unit

6 Chemical Energy

In your life you encounter many different kinds of energy: radiant energy from our sun warms you on a spring day; electrical energy provides the energy to run your computer; chemical energy from food keeps your body functioning. In each of these cases energy is transformed from one form to another.

Energy transformations are the basis for all change, everywhere. Photosynthesis is a chemical process that takes radiant energy from the sun and stores it within molecules of glucose. This stored chemical energy provides the energy for life. Fossil fuels also contain stored chemical energy that can be transformed into other forms of energy. Energy links together all components of the universe as the energy is emitted, captured, stored, and used. All physical, chemical, and nuclear changes involve changes in energy, but this unit will focus on chemical changes involving energy: thermochemical changes, or *thermochemistry.*

Consider the Alberta photo on the right. Solar energy from the nuclear reactions in our sun is being used through photosynthesis by the barley plants. Carbohydrates are being chemically produced, which may be used to fuel us as human beings. The pumpjack is removing conventional crude oil from below the surface. The oil, we believe, was produced from the decay of plants. The oil is then processed by chemical technologies to, for example, run our cars and trucks. Solar energy, directly or indirectly, is the source of most energy available to us on Earth.

Everything humans do requires energy. It is a major factor in social change on our planet. Technologies, which inevitably consume energy, are developed for a social purpose. These technologies, however, often have drawbacks related to their use of energy. The control and use of our present sources of energy, as well as the development of new sources, will continue to have far-reaching environmental, economic, social, technological, and political implications for many years.

As you progress through the unit, think about these focusing questions:

- How does our society use the energy of chemical changes?
- How do chemists determine how much energy will be produced or absorbed for a given chemical reaction?

GENERAL OUTCOMES

In this unit, you will

- determine and interpret energy changes in chemical reactions
- explain and communicate energy changes in chemical reactions

Unit 6 Chemical Energy

ARE YOU READY?

These questions will help you find out what you already know, and what you need to review, before you continue with this unit.

Prerequisites

Concepts

- sources of energy (e.g., solar and chemical)
- kinetic and potential energy
- thermal energy
- temperature and phase changes
- heat
- transfer of energy
- conservation of energy
- kinetic molecular theory
- photosynthesis
- cellular respiration
- carbon cycle
- endothermic and exothermic reactions

Skills

- measurement
- graphing

You can review prerequisite concepts and skills on the Nelson Web site, in the Chemistry Review unit, and in the Appendices.

A Unit Pre-Test is also available online.

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Knowledge

1. Most of the energy available on Earth comes from our sun. For each of the Alberta-based technological devices in **Figure 1**, briefly trace the energy produced back to its source, our sun, Sol.

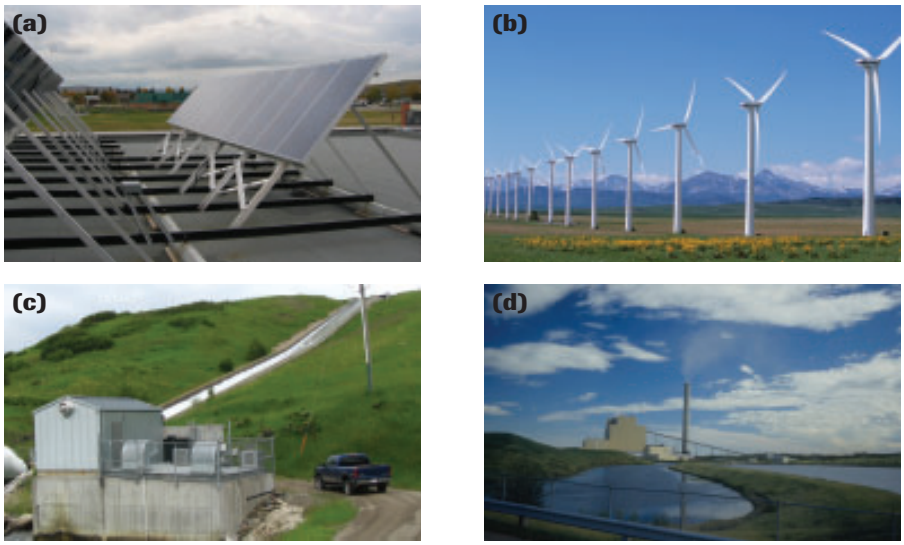


Figure 1
(a) photovoltaic (solar) cells: Students and teachers have installed photovoltaic technology at Cochrane High School, Alberta.
(b) wind turbine: The Castle River Wind Farm is one of many in Alberta, which (along with Quebec) leads Canada in wind-generated electricity.
(c) hydroelectric power plant: Belly River Hydroelectric Plant is a low-impact generating station near Glenwood, Alberta.
(d) coal-fired power plant: Sundance Powerplant is on the south shore of Lake Wabamun, Alberta.

2. Write balanced chemical equations for the complete combustion of the following fuels:
   (a) coal (assume C(s))
   (b) propane ($C_3H_8(g)$)
   (c) gasoline (assume $C_8H_{18}(l)$)

3. List five energy sources that do not directly produce carbon dioxide.

4. The overall reaction for photosynthesis can be written as
   $6\ CO_2(g) + 6\ H_2O(l) + \text{sunlight} \rightarrow C_6H_{12}O_6(aq) + 6\ O_2(g)$
   (a) What additional substance must be present for this reaction to occur?
   (b) Is photosynthesis an endothermic or exothermic process?
   (c) What industries depend heavily on the chemistry of photosynthesis?

5. The overall reaction for cellular respiration can be written as
   $C_6H_{12}O_6(aq) + 6\ O_2(g) \rightarrow 6\ CO_2(g) + 6\ H_2O(l)$
   (a) Cellular respiration is similar to what common type of reaction?
   (b) If energy were to be included as a term in this reaction equation, should it be written on the right or left side of the arrow?

(c) Is cellular respiration an endothermic or exothermic process?
(d) Give two examples of cellular respiration.

6. Explain why fossil fuel energy can be considered as stored solar energy.

7. The carbon cycle involves the interrelationship of living things with the environment and with technologies that use fossil fuels. Draw a diagram of the carbon cycle using examples from the natural world, (such as plants and animals) and from the technological world (such as transportation and petrochemicals). Use arrows to illustrate the relationships (by the processes of cellular respiration, photosynthesis, or combustion) among these aspects of life in Alberta.

8. From a theoretical perspective, kinetic energy is related to the motion of an entity. The *temperature* of an object is a measure of the average kinetic energy of its entities, while the *thermal energy* of an object is the total kinetic energy of its entities. *Heat* is energy being transferred between two objects or materials (**Figure 2**). What happens to the average kinetic energy of water molecules when the water is
(a) heated
(b) cooled
(c) maintained at a constant temperature
(d) boiled

9. Potential energy is a stored form of energy. Chemical potential energy involves the energy stored in the bonds between atoms and molecules. State what happens to the potential energy of water molecules when the water is
(a) heated
(b) boiled
(c) frozen

10. The motion of molecules can be translational (straight-line), rotational, and/or vibrational. Describe the dominant types of motion of water molecules in
(a) ice
(b) liquid water
(c) water vapour

11. A pot of water at room temperature is placed on the burner of a gas stove and heated until 15% of the water has evaporated due to boiling.
(a) Copy **Figure 3**, and complete it to show a temperature–time graph for the heating (including boiling) of the water.
(b) Describe the changes that take place in the kinetic and potential energy of the water molecules as the water is heated, and then when it boils.
(c) Describe the changes that take place in the kinetic and potential energy of the gas molecules as they undergo combustion in the stove.
(d) Use the law of conservation of energy to explain the graph that you have drawn.

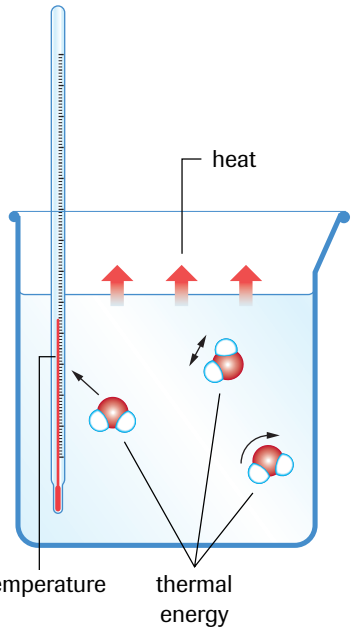


Figure 2
A thermometer (or a temperature probe) measures the average kinetic energy of the molecules in a beaker of water.

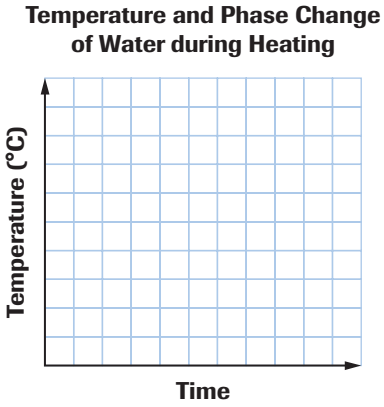


Figure 3
Temperature–time axes

STS Connections

12. List three energy sources that are renewable and three that are nonrenewable.

13. What is your reasoned response to someone trying to sell you shares in a corporation that is manufacturing a perpetual motion machine?

Skills

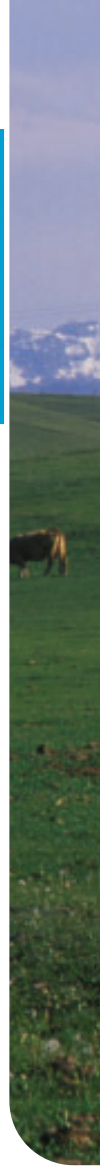
14. Laboratory thermometers are usually read to ±0.2 °C.
(a) Interpret a thermometer reading of 22.4 °C on this thermometer.
(b) Critique the communication when someone interprets a reading of 22.0 °C as "exactly twenty-two degrees Celsius" and records the temperature as 22 °C.

chapter

11 Enthalpy Change

In this chapter

- Exploration: Burning Oil
- Case Study: Personal Use of Chemical Energy
- Investigation 11.1: Designing and Evaluating a Calorimeter
- Web Activity: Thermal Insulation
- Lab Exercise 11.A: Molar Enthalpy of Neutralization
- Web Activity: Calorimetry
- Investigation 11.2: Molar Enthalpy of Reaction
- Web Activity: Hess' Law
- Lab Exercise 11.B: Testing Hess' Law
- Lab Exercise 11.C: Analysis Using Hess' Law
- Investigation 11.3: Applying Hess' Law
- Explore an Issue: Alternative Energy Sources and Technologies
- Lab Exercise 11.D: Testing $\Delta_f H°$ from Formation Data
- Web Activity: Rocket Fuel Thermochemistry



We have invented many technologies to use energy: campfires and wood stoves use energy stored in wood; stoves, heaters, and cars used energy stored in liquid hydrocarbons; industries use energy stored in natural gas; and power plants convert the chemical energy stored in natural gas, oil, and coal into electrical energy. Trucks, trains, and airplanes use energy stored in fuels refined from crude oil, heavy oil, and oil sands.

The future could look quite different from the present. Transportation technologies may derive their energy from the combustion of alternative fuels such as hydrogen, ethanol, and methanol. Hybrid cars, operating on both gasoline and batteries, are already becoming more common. Science and technology work in parallel with one another—although sometimes one is ahead of the other—often driven by social pressures. Sometimes a scientific concept is applied by technology; other times science is used to explain a technology after it is invented (independent of the science).

Consumers will have many choices to make about energy technologies. Knowing the science and technology of energy is advantageous. Energy-related issues are important from many perspectives. Will you support politicians who advocate further development of Alberta's oil sands? Will your next car be gasoline powered? This chapter will help you make these decisions in a more informed manner.

STARTING Points

Answer these questions as best you can with your current knowledge. Then, using the concepts and skills you have learned, you will revise your answers at the end of the chapter.

1. Consider the following changes: photosynthesis, cellular respiration, and combustion of gasoline. Classify each of these changes as exothermic or endothermic.
2. How is electrical energy produced in Alberta? What are the sources of energy that produce this electricity?
3. Select one energy source for producing electrical energy, and list three benefits and three risks.
4. Trace the transformations of energy from solar radiation through to using a fluorescent light bulb at home.
5. (a) How can we measure how much energy is released or absorbed by a chemical reaction?
   (b) How can we communicate how much energy is released or absorbed by a chemical reaction?

Career Connection: Chemical Engineering Technologist

Figure 1
Our lives depend on energy transformations: from solar energy to wind energy to electrical energy; and from solar energy to chemical energy (food for cattle and ourselves).

Exploration Burning Oil

Engineers who design furnaces to heat homes and nutritionists who calculate the energy value of different foods need to analyze the energy supplied by different fuels. In experiments in which thermal energy is absorbed by water, they use the equation

$$Q = mc\Delta t$$

(thermal energy = mass × specific heat capacity × temperature change)

($c = 4.19$ J/(g·°C) for water)

The general concept of what these technology experts do is illustrated by the following low-technology experiment.

Materials: eye protection; lab apron; centigram balance; unshelled pecan or other nut; paper clip; small tin can, open at one end and punctured under the rim on opposite sides; pencil; thermometer; beaker; matches; laboratory stand with ring

- Measure the mass of the nut or assume that an average pecan has a mass of about 0.5 g.
- Bend a paper clip so that it forms a stand that will support a nut above the lab bench (**Figure 2**).
- Place 50 mL of tap water in the tin can. Measure the temperature of the water.
- Suspend the can of water inside the ring and above the nut by putting the pencil through the holes and under the rim of the can.
- Using a match, light the nut. When it has finished burning, measure the final temperature of the water.

Figure 2
Experimental set-up

(a) Calculate how much energy was absorbed by the water.
(b) Where did this energy come from?
(c) Calculate the quantity of energy produced per gram of fuel (nut) burned.
(d) Compare this combustion reaction to the reaction that would happen if you were to eat the nut instead of burning it.
(e) This experiment uses low-tech equipment. How can you improve on the technology and, therefore, improve on the certainty of the experimental results?

Students with sensitivity to nuts or nut products should not perform this activity.

11.1 Energy Demands and Sources

Photosynthesis is a major natural contributor to stored chemical energy on Earth. Research has indicated that photosynthesis takes energy from our sun, Sol, and converts this radiant (electromagnetic) energy into chemical energy (**Figure 1**). Carbon dioxide and water react in plants, in the presence of sunlight, to produce glucose and oxygen. This chemical (potential) energy is consumed by animals, such as ourselves, and is stored or used to fuel life. The chemical energy stored by photosynthesis is accessed through cellular respiration, in which glucose and oxygen react to produce carbon dioxide and water and to supply energy. (This overall reaction is the reverse of photosynthesis.)

As presented in the carbon cycle in Chapter 9, the most accepted hypothesis for the origin of fossil fuels is the biogenic hypothesis. This hypothesis suggests that fossil fuels had their origins in deposits of plant and animal material. Solar energy was captured and stored by photosynthesis. The plant matter was either eaten by animals and stored as chemical potential energy (in carbohydrates, protein, and fat) in the body of the animal before the animal died, or the plant itself died. In certain geological areas, time, temperature, and pressure converted the dead animal and plant material, with its stored chemical potential energy, into hydrocarbons. This line of reasoning is used to support the claim that fossil-fuel energy has its origins in solar energy from our sun.

Fossil fuels are a major source of stored chemical energy. Humans have invented technologies to extract and process fossil fuels (**Figure 1**). In Alberta, huge technological advances have been made for recovering chemical energy from, in chronological order, coal, natural gas, crude oil, heavy oil, oil sands, and coal-bed methane. The demand for energy has challenged the technological inventiveness of engineers, scientists, and others in the energy workplace. In Alberta, for example, nearly 20% of all jobs are in the energy sector.

DID YOU KNOW ?

Energy Production and Consumption

In the 1960s Canada moved from being a net consumer of energy to a net producer of energy. By 2002 Canada was producing 45% more energy than it consumed. In the same year, according to Statistics Canada, Canadian energy consumption per capita reached a record high. More recent statistics may, however, rewrite that record.

CAREER CONNECTION

Chemical Engineering Technologist

This career requires problem-solving skills and creativity, and there are many different options for jobs. Chemical engineering technologists may work in production facilities, where they operate process control equipment and supervise construction activity, or they may work in engineering offices, where they redesign process equipment and perform computer-based simulations. Explore the variety of tasks performed in this career, and learn what training is needed.

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Figure 1
Crude oil, pumped near Longview, Alberta, is an indirect source of solar energy.

Transportation and building designers have to use just as much ingenuity as the engineers and scientists working in the energy sector. On the production side, the processes have been made more efficient and economical. On the consumption side, the technologies (such as cars and homes) also need to be more efficient and economical. Of course, there are other perspectives, such as environmental, legal, ethical, and social, that also need to be considered when designing a technology for converting and using energy.

Chemists have created some concepts that help us explain chemical technologies and the transformation of energy to and from chemical energy. These concepts are conveyed using chemical energy terms, some of which are presented in the Summary on page 493. These definitions are simplified. If you take more in-depth chemistry courses you will learn more elaborate definitions, but these will suit us very nicely for now.

There are four major demands for energy from fossil fuels—heating, transportation, industry, and commercial and institutional. What are some alternatives to using fossil fuels? Options include both the use of different fuels and more economical management of fossil fuels (**Table 1**).

Table 1 Alternatives to Current Fossil Fuel Uses

Energy demands	Alternative energy sources and practices
heating residential (~15%)	• solar heating (**Figure 2**), heat pumps, geothermal energy, biomass gas, and electricity from hydroelectric and nuclear plants • improved building insulation and design
transportation (~30%)	• alcohol/gasohol and hydrogen fuels (**Figure 3**), and electric and hybrid (electric and gasoline) vehicles • mass transit, bicycles, and walking
industry (~40%)	• solar energy, nuclear energy, and hydroelectricity • improved efficiency and waste heat recovery (**Figure 4**)
commercial and institutional (~15%)	• solar water pre-heating and heat pumps, including geothermal • water and heat conservation

DID YOU KNOW ?

Electric Cars

Although electric cars are making a comeback in popularity, at one time they were more prevalent than gasoline-powered cars. For example, in 1900, in the cities of New York, New York, Chicago, Illinois, and Boston, Massachusetts, there was a total of 2370 automobiles: 400 gasoline-powered, 800 battery-powered, and 1170 steam-driven.

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DID YOU KNOW ?

Living Lightly on the Earth

Many Aboriginal lifestyle philosophies embrace the idea of taking as little as possible from the environment and leaving more for future generations. In the 21st century, we are using energy at an accelerating rate, especially nonrenewable energy in the form of fossil fuels. To live more lightly on the Earth should we be looking for a technological fix or personal lifestyle changes?

Figure 2
A well-insulated home with the majority of windows facing south is the main requirement for obtaining heat from direct sunlight. Solar heating and heat generated by people and appliances can reduce heating bills by as much as 90%. (A model home in the Drake Landing Solar Community, Okotoks, Alberta.)

Figure 3
Fuel cell vehicles, such as the DaimlerChrysler F-Cell pictured here, are powered by Ballard® fuel cell technology and use hydrogen as a fuel. Fuel cell vehicles emit only heat and water vapour, making them true zero-emission vehicles.

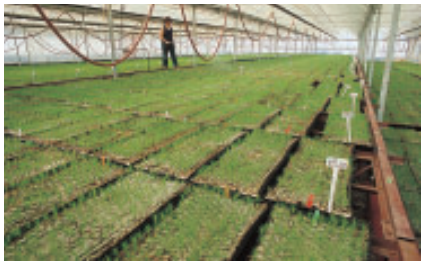

Figure 4
A great deal of energy is wasted by motors, compressors, and exhaust emissions. Recovering this energy is one way in which industries improve their energy efficiency. This greenhouse uses heat from the compressors on the TransCanada Pipeline.

Personal Use of Chemical Energy

We rely on chemical energy for our very lives, as well as to heat our homes, cook our food, and move us around. Chemical energy for personal use comes from two main sources: combustion reactions and electrochemical reactions. The most personal use of chemical energy is eating food. The human body converts the chemical energy in food into mechanical energy and body heat through the process of cellular respiration, a form of combustion. A typical high-calorie fast-food meal (burger, fries, and drink) provides about 6 MJ of energy. (Research indicates that daily intakes would be around 9 to 12 MJ for a normally active young adult, and up to 29 MJ for an extreme high-performance athlete.) When you buy food, you are buying energy to keep your body running.

First Nations peoples of Alberta traditionally consumed much less energy in their balanced, self-sustaining mode of life. It has been generally noted by anthropologists that earlier lifestyles, barring accidents and foreign disease, were much more healthful for individuals than current lifestyles, which too often include diets high in fat and sugar.

In Alberta, nearly all the energy for home heating and cooking comes from the combustion of fossil fuels, either directly or indirectly. For example, many homes are heated directly by burning methane or propane on-site, or by electricity, which may be produced by burning fossil fuels off-site. Propane and butane are convenient fuels for motor homes and campers because these hydrocarbons can easily be compressed, stored, and transported. A century ago in Alberta, firewood and coal were important fuels for heating and cooking, but now firewood is used mainly by wilderness campers and people living in remote areas.

Our transportation needs are also supplied by chemical energy (**Figures 5** and **6**). The family car and motorcycle use gasoline, which stores a large quantity of energy in a small volume. The combustion of a litre of gasoline produces about 34 MJ of energy.

Figure 5
Cycling uses food energy.

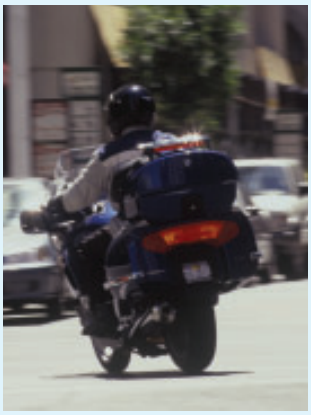

Figure 6
Motorcycling uses fossil fuel energy.

Gasohol is a mixture of gasoline and methanol/ethanol, designed to reduce pollution and our reliance on fossil fuels. Renewable fuels derived from biomass (such as wood pulp, grain, and crop by-products) are used together with fossil fuels.

The technology for using hydrogen gas to run automobiles has been developed but its use is not widespread. Chemical energy stored in batteries is used to power golf carts, passenger cars, and delivery vans. A battery-powered vehicle can travel only a relatively short distance before the battery needs to be recharged. Hybrid automobiles combine a gasoline engine with a battery-powered electric motor that becomes a generator and recharges as you drive. Hybrid vehicles are also able to use gravitational potential energy when going downhill, and kinetic energy when braking, to help recharge the battery.

The portability of small batteries makes them valuable for radios, cameras, music players, phones, and personal computers.

Most of our personal use of chemical energy, however, comes from burning fossil fuels—even the energy used to recharge batteries likely comes from an electrical energy plant that burns coal or natural gas. Our dependence on fossil fuels raises concerns about depleting our energy reserves and increasing greenhouse gas emissions. On a personal level, a simple solution is to drive a car less and walk, bike, or ride a bus more often. The chemical energy for self propulsion comes from food, and the carbon dioxide produced by cellular respiration is significantly less than that produced by a car.

Case Study Questions

1. What are the two main sources of chemical energy for personal use?
2. When 1 mol of glucose is burned, 2802.5 kJ of energy is released. Calculate the quantity of energy released to a person by eating 5.00 g of glucose in a candy.
3. How could technology help us meet our need for energy? Suggest at least four ways.
4. Technological solutions (fixes) are often chosen over changing personal habits or behaviour. Some people argue that saving fossil-fuel energy for future generations is important. Provide three examples of technological fixes versus changes in personal behaviour for reducing chemical energy consumption. Briefly evaluate, for each example, which would be the better solution.
5. Identify and list the perspectives represented in this Case Study. See Appendix D.1.

Section 11.1 Questions

Use **Table 2** to answer questions 1 and 2. Alternatively, go to the Statistics Canada Web site and obtain the latest data.

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Table 2 Canadian Energy Production by Fuel Type (Source) in 2003 (in terajoules (TJ), or 10^{12} J)

Coal (TJ)	Crude oil (TJ)	Natural gas (TJ)	Natural gas liquids (NGLs) (TJ)	Primary electricity, hydro and nuclear (TJ)	Total production (TJ)
1 326 114	5 679 573	7 024 602	642 897	1 457 467	16 130 653

Source: Statistics Canada

1. (a) Convert the Canadian energy production into a percentage of the total for each fuel type for the latest year provided. For consistency of use in the pie chart in part (b), give each percentage to one decimal place.
(b) Display these percentages as a pie chart. Use graphing software if available.

2. (a) Go to the Statistics Canada Web site and calculate the percentage of the production of each fuel type that is exported from Canada.
(b) Display these percentages as a bar graph. Use graphing software if available.

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Use **Table 3** to answer questions 3 to 6. Alternatively, go to the Statistics Canada Web site and obtain the latest data.

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Table 3 Canadian Energy Use/Demand (in terajoules (TJ), or 10^{12} J)

Energy Use/Demand by Year	1999	2000	2001	2002	2003
Total industrial (TJ)	2 177 297	2 268 624	2 166 287	2 229 541	2 313 106
Total transportation (TJ)	2 307 283	2 279 845	2 240 367	2 250 130	2 242 042
Agriculture (TJ)	229 865	231 927	218 075	206 753	211 866
Residential (TJ)	1 232 263	1 287 825	1 239 970	1 286 677	1 348 041
Public administration (TJ)	124 522	131 288	126 813	125 164	127 678
Commercial and other institutional (TJ)	1 061 446	1 176 423	1 184 065	1 286 657	1 362 441
Total energy use (TJ)	**7 132 504**	**7 375 967**	**7 175 442**	**7 384 682**	**7 604 948**

Source: Statistics Canada

3. For the latest year listed, what are the percentages of use (to one decimal place) by the sectors listed?

4. Create a line graph for the total energy use (in petajoules (PJ; $\times 10^{15}$ J)) over the years for which data are provided. Use graphing software if available.

5. Create a line graph for residential use (in terajoules (TJ)) over the years for which data are provided. Use graphing software if available.

6. Go to the Statistics Canada Web site and obtain the breakdown of residential use by fuel type.
(a) What percentage of residential energy use (to one decimal place) is from natural gas?
(b) Quote three other statistics that you can extract from the data. Be sure to cite your sources.

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7. There is an increasing number of energy choices for operating a car. List at least four of these choices.

8. Homes in Alberta are heated by a variety of energy sources. List at least four of these sources.

9. The Drake Landing Solar Community in Okotoks, Alberta, is an experimental housing project that uses solar energy to heat the homes. Run the animation through the Nelson Web site, and write a short description of the project from a technological perspective.

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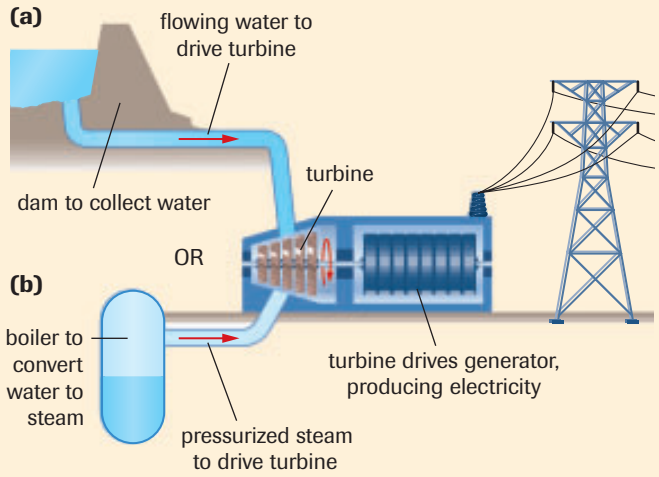
10. Energy consumption for keeping our homes warm (and supplied with hot water) varies fairly predictably during the year (**Table 4**).

Table 4 Average Monthly Energy Use in Alberta per Household

Month	Energy (GJ)
May	5
June	3
July	3
August	4
September	5
October	9
November	18
December	22
January	23
February	18
March	16
April	9
Annual total	**135**

(a) Prepare a bar graph to express the monthly data. Use graphing software if available.
(b) From energy bills, how does your household energy use compare with these values?

11. Several energy sources are available for heating: oil, gas, coal, wood, solar, geothermal, and nuclear. These sources produce heat directly. Indirect sources of heat include electric motors and lights.
(a) Which of the direct energy sources originated in our sun, and which originated from the formation of our planet?
(b) Which two energy sources are currently, in your area, most commonly used for heating?
(c) Which two energy sources do you think will be the most widely used for heating by the year 2030?



12. Engineers and scientists have invented technologies for using locally available sources of energy to produce electricity. In **Figure 7**, the kinetic energy of falling water or of steam is used to turn a turbine attached to an electric generator.
(a) What fuels can be burned to produce the steam?
(b) What other sources of energy (other than hydro and fossil fuels) are available to produce electricity?

13. Some retailers have advertised titanium necklaces as an energy source. As evidence to support their claim, they ask customers to lift a bag containing a brick. They then drape the necklace over the customer's arm and have the person lift the bag again. Customers find that lifting the bag with the necklace present is easier, and they often buy the necklace as a result. Critique this experimental design.

Extension

14. Energy consumers were briefly excited when successful "cold fusion" was announced by two scientists in 1989. Successful cold fusion would represent an inexpensive, clean, readily available source of energy. Do some library research or search the Internet to learn more about the cold fusion experiments.
(a) Were the two scientists empiricists or theoreticians?
(b) Were they qualified to make theoretical interpretations of their experimental work?
(c) Are there cold-fusion research programs still being funded? If so, where?
(d) Evaluate your literature research strategies.

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15. The caption for Figure 3 (page 481) says that these cars are "true zero-emission vehicles." Consider this claim. Research the environmental effects of water vapour. Think about the research and production involved in developing such vehicles. Write or present a brief rebuttal or supporting statement for the "zero-emissions" claim.

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Figure 7
There are many energy sources that can be used to spin a turbine. If the turbine is connected to a generator, the energy of motion is converted to electrical energy.

Calorimetry 11.2

According to the law of conservation of energy, energy is neither created nor destroyed in any physical or chemical change. In other words, energy is only converted from one form to another. The study of energy changes (that is, energy produced or absorbed) by a chemical system during a chemical reaction is called **thermochemistry**. To study energy changes chemists require an **isolated system**, that is, one in which neither matter nor energy can move in or out. They also need carefully designed experiments and precise measurements. **Calorimetry** is the technological process of measuring energy changes of an isolated system called a **calorimeter** (**Figure 1**). Within this isolated system the chemical system (reactants and products) being studied is surrounded by a known quantity of liquid (generally water) inside the calorimeter. Energy, as **heat**, is transferred between the chemical system and the water. (Note that heat as a form of energy can only be transferred between substances; heat is never possessed by them.) The energy gained by the chemical system is equal to the energy lost by the calorimeter and its contents, as long as both the chemical system and the calorimeter (the surroundings) are part of an isolated system. In other words, for the measurement to be accurate, no energy may be transferred between the inside of the calorimeter and the environment outside the calorimeter.

The main assumption is that no heat is transferred between the calorimeter and the outside environment. Although chemists know this to be a false assumption, it provides a close approximation. The better the calorimeter technology, the closer the assumption is to being valid. A simplifying assumption is that any heat absorbed or released by the calorimeter materials, such as the container, is negligible. Another assumption is that a dilute aqueous solution in a calorimeter has the same physical properties as pure water (for example, the same volume and specific heat capacity).

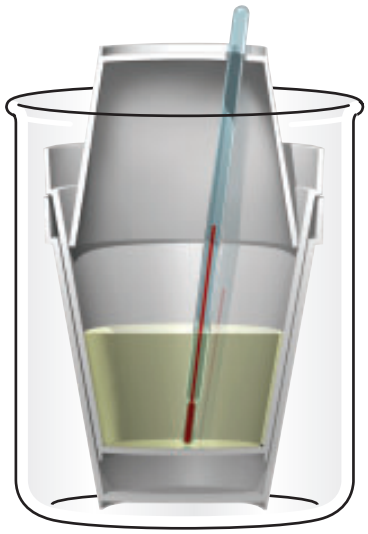

Figure 1
A simple laboratory calorimeter consists of an insulated container made of three nested polystyrene cups, a measured quantity of water, and a thermometer. The chemical system is placed in or dissolved in the water of the calorimeter. Energy transfers between the chemical system and the surrounding water are monitored by measuring changes in the water temperature.

Analyzing Energy Changes

When methane reacts with oxygen in a laboratory burner, enough energy is transferred to the surroundings to increase the temperature (**Figure 2**). How is this energy quantified? Calorimetry requires careful measurements of masses and temperature changes. When a fuel such as methane burns, energy is released and heat is transferred from the chemical system to the surroundings (which include the water in the beaker). Given the same quantity of heat, a small quantity of water will undergo a greater increase in temperature than a large quantity of water. In addition, different substances vary in their ability to absorb quantities of thermal energy (**Figure 3**). Finally, if more heat is transferred, the observed temperature rise in the water is greater.

Calorimetry work in the laboratory has shown that, keeping the type of substance constant (for example, using only water in the calorimeter), the quantity of energy transferred is directly proportional to the mass of the substance and its change in temperature. The chemical energy lost by the chemical system is calculated from the thermal energy gained by the surroundings. **Thermal energy** is the total kinetic energy of the entities of a substance (for example, the molecules of water in a calorimeter).

These three factors—mass (m), type of substance, and temperature change (Δt)—are combined in an equation to represent the quantity of thermal energy (Q) transferred:

$$Q = mc\Delta t$$

where c is the **specific heat capacity**, the quantity of energy required to raise the temperature of a unit mass (for example, one gram) of a substance by one degree Celsius or one kelvin. For example, an average specific heat capacity of water around room temperature is 4.19 J/(g·°C). (Recall that the SI unit for energy is the **joule**, J.)

Figure 2
The fuel in the burner releases energy that is absorbed by the surroundings, which include the beaker, water, and air.

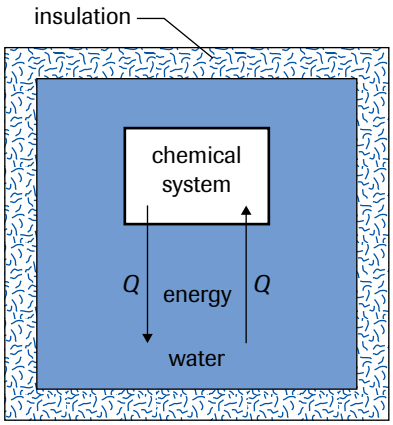

Figure 3
The chemical system inside the calorimeter undergoes either a phase change, such as fusion, or a chemical change, such as a double replacement reaction. Energy is either absorbed from the water or released to the water. An increase in the temperature of the water indicates an exothermic change of the system; a decrease in the temperature of the water indicates an endothermic change of the chemical system.

Table 1 Specific Heat Capacities of Substances

Substance	Specific heat capacity, c (J/(g·°C))
ice	2.00
water	4.19
steam	2.02
aluminium	0.897
iron	0.449

Specific heat capacities vary from substance to substance, and even for different states of the same substance (**Table 1**). The specific and volumetric heat capacities listed in Appendix I were determined by means of calorimetry.

COMMUNICATION example 1

Determine the change in thermal energy when 115 mL of water is heated from 19.6 °C to 98.8 °C.

Solution

$$Q = mc\Delta t$$
$$= 115\text{ g} \times \frac{4.19\text{ J}}{\text{g·°C}} \times (98.8 - 19.6)\text{ °C}$$
$$= 38.2\text{ kJ}$$

Based upon the assumption that the temperature change is a complete measure of the change in thermal energy, the change in thermal energy of the water is 38.2 kJ.

Some of the calorimetry assumptions are that

- all the energy lost or gained by the chemical system is gained or lost (respectively) by the calorimeter; that is, the total system is isolated.
- all the material of the system is conserved; that is, the total system is isolated.
- the specific heat capacity of water over the temperature range is 4.19 J/(g·°C).
- the specific heat capacity of dilute aqueous solutions is 4.19 J/(g·°C).
- the density of a dilute aqueous solution is the same as that of water; that is, 1.00 g/mL. (see Communication Example 1)
- the thermal energy gained or lost by the rest of the calorimeter (other than water) is negligible; that is, the container, lid, thermometer, and stirrer do not gain or lose thermal energy.

DID YOU KNOW?

A Kilojoule of Energy

Completely burning a wooden match releases about 1 kJ of energy. This quantity of energy could heat $\frac{1}{4}$ cup of water by about 1 °C, and is equivalent to about a quarter of a calorie of food energy. There are two definitions of calorie but only one for kilojoule: 1000 J.

EXTENSION

Food Energy

Labels on food packaging report the number of kilojoules per serving of the contents. How do scientists determine this number? This audio clip will help you understand how thermochemistry is used to determine the quantity of energy available in our food.

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INVESTIGATION 11.1 *Introduction*

Designing and Evaluating a Calorimeter

A calorimeter can be as simple as an uninsulated beaker or can. Literature research indicates that very sophisticated calorimeters, such as bomb and flow calorimeters, are also available. There are many laboratory calorimeters between these extremes. In this investigation you will create your own calorimeter. Use modern materials and/or consider what Aboriginal peoples might have used for insulating a calorimeter-like vessel.

Purpose

The purpose of this investigation is to test a hypothetical design of a calorimeter that is improved (based upon your own scientific and technological criteria) over the simplest calorimeter.

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
● Hypothesis	● Procedure	
○ Prediction	● Evidence	

Problem

What is the best design for a simple calorimeter?

Design

Each research group in class uses a different calorimeter design of their choosing. One group uses an uninsulated calorimeter as a control. An individual group and a class Analysis are employed.

To perform this investigation, turn to page 516.

Practice

1. In your own words, write the law of conservation of energy.
2. Why do we call a calorimeter an "isolated system"?
3. A calorimeter is used to measure (somewhat indirectly) what kind of energy?
4. Experiments have shown that a thermal energy change is affected by mass, specific heat capacity, and change in temperature. What happens to the thermal energy if
   (a) the mass is doubled?
   (b) the specific heat capacity is divided by two?
   (c) the change in temperature is tripled?
   (d) all three variables are doubled?
5. If the same quantity of energy were added to individual 100 g samples of water, aluminium, and iron, which substance would undergo the greatest temperature change? Explain. (See Appendix I for specific heat capacities.)
6. After coming in from outside, a student makes a cup of instant hot chocolate by heating water in a microwave. What is the gain in thermal energy if a cup (250 mL) of tap water is increased in temperature from 15 °C to 95 °C?
7. A backpacker uses an uncovered pot to heat 1.0 L of lake water on a single-burner stove. If the water temperature rises from 5.0 °C to 97.0 °C, what is the gain in thermal energy of the water?
8. To test the efficiency of an energy transfer, a chemist supplies 30.0 kJ of external energy to a simple calorimeter. If 150 mL of water in the calorimeter is heated from 20.6 °C to 52.8 °C, what is the percent efficiency of the calorimeter?

Heat Transfer and Enthalpy Change

Chemical systems have many different forms of energy, both kinetic and potential. These include the kinetic energies (energy of motions) of

- moving electrons within atoms
- the vibration of atoms connected by chemical bonds
- the rotation and translation of molecules that are made up of these atoms.

The chemical potential energy of a chemical system includes energy stored in

- covalent and/or ionic bonds between the entities
- intermolecular forces between entities

DID YOU KNOW?

Honeymoon Inspiration

During his honeymoon, while sitting beside a waterfall, James Joule speculated on the temperature increase of the water as it fell over the waterfall. Joule tenaciously pursued the relationship between work and heat, testing his hypothesis that the work needed to produce a certain quantity of heat is always the same. This realization led to the idea that heat is a form of energy rather than a fluid, as scientists believed at the time.

Joule is known for his patient and precise quantitative measurements that allowed him to calculate the specific heat capacity of water. He also studied the relationship between heat and temperature and helped establish the law of conservation of energy.

Figure 4
James Prescott Joule (1818–1889)

Learning Tip

In searching through references you may find the terms *enthalpy of reaction*, *heat of reaction*, *enthalpy change*, and ΔH. You can assume, at this point, that they all mean the same thing.

The total of the kinetic and potential energy within a chemical system is called its enthalpy. The enthalpy of a chemical system is an expression of the chemical energy possessed by the chemical system. However, if the reactants and products are at the same initial and final temperature, then the kinetic energy of the system does not change.

Just as electrical and gravitational potential difference are most often communicated as a difference from a reference point, enthalpy is most often communicated as a difference in enthalpy between reactants and products for a particular chemical system. An **enthalpy change**, ΔH, communicates the difference between the enthalpy (assume the difference in chemical potential energy) of the products and the enthalpy of the reactants.

$$\Delta H = H_{\text{products}} - H_{\text{reactants}} \text{ or } \Delta H = H_{\text{P}} - H_{\text{R}} \text{ (IUPAC definition)}$$

DID YOU KNOW

Setting Hard

The setting of concrete is quite exothermic, and the rate at which it sets or cures determines the hardness of the concrete. If the concrete sets too quickly (for example, if the enthalpy of reaction is not dissipated quickly enough into the air), the concrete may expand and crack. Doctors are cautious of similar heating when making a plaster cast for a broken limb.

Analysis of calorimetric evidence is based on the law of conservation of energy and on several assumptions. The law of conservation of energy may be expressed in several ways, for example, "The total energy change of the chemical system is equal to the total energy change of the calorimeter surroundings." Using this method, both the enthalpy change, ΔH, and the quantity of thermal energy, Q, are calculated as absolute values, without a positive or negative sign.

$$\underset{\text{(system)}}{\Delta H} = \underset{\text{(calorimeter)}}{Q}$$

Energy loss by the chemical system is equal to energy gain by the surroundings. Consider the reaction that occurs when zinc metal is added to hydrochloric acid in an open flask:

$$Zn(s) + 2\ HCl(aq) \rightarrow H_2(g) + ZnCl_2(aq)$$

Research on this reaction indicates that the solution increases in temperature. In other words, there is an increase in thermal energy of the surroundings. Chemists infer from the temperature increase that the chemical system (in this case, Zn(s) and HCl(aq)) must have lost chemical energy to the surroundings (**Figure 5**). At this stage in your understanding, you can make the assumptions that $\Delta E_{\text{p}} = \Delta H$ and $\Delta E_{\text{k}} = Q$.

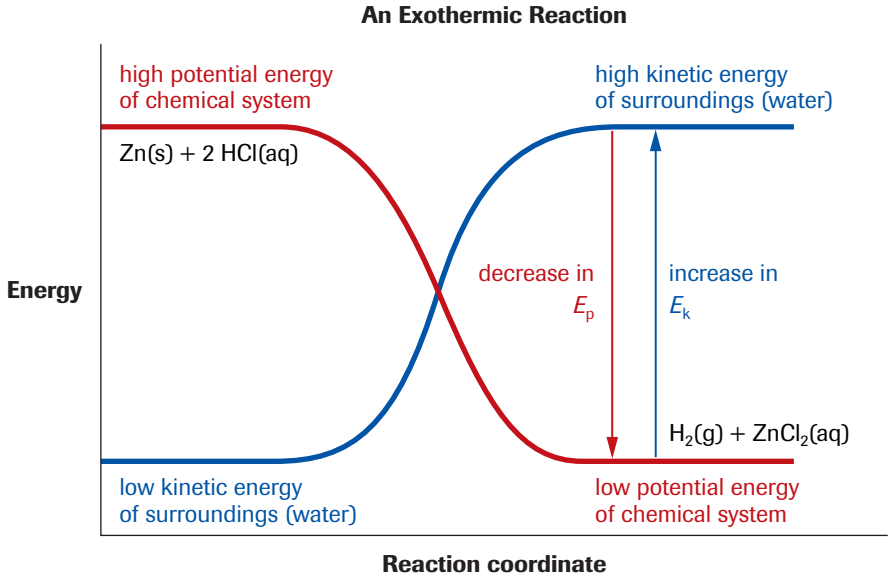

Figure 5
In this example of an exothermic change, the change in potential energy of the chemical system (ΔE_{p} or ΔH) equals the change in kinetic energy of the surroundings (ΔE_{k} or Q). This is consistent with the law of conservation of energy.

SAMPLE problem 11.1

In a simple calorimetry experiment involving a burning candle and a can of water, the temperature of 100 mL of water increases from 16.4 °C to 25.2 °C when the candle is burned for several minutes. What is the enthalpy change?

The first step in solving this problem is to use the law of conservation of energy. According to this law, the energy lost by the chemical system (the burning candle) is equal to the energy gained by the surroundings (assume the calorimeter water). The change in enthalpy ($\Delta_c H$) of the chemical system can be determined from the change in thermal energy (Q) of the surroundings, as expressed by the equation

$$\Delta_c H = Q$$

Since $Q = mc\Delta t$

$$\Delta_c H = mc\Delta t$$

Substituting the evidence gathered into the conservation of energy equation, and using the density of water as 1.00 g/mL

$$\begin{aligned}\Delta_c H &= mc\Delta t \\ &= 100\text{ g} \times \frac{4.19\text{ J}}{\text{g}\cdot^\circ\text{C}} \times (25.2 - 16.4)\,^\circ\text{C} \\ &= 3.69\text{ kJ}\end{aligned}$$

Based upon the evidence gathered in this experiment, the approximate value for the enthalpy change of the chemical system is *recorded* as −3.69 kJ.

Learning Tip

The following symbols are used by IUPAC to denote a chemical process or reaction.

Table 2 Symbols for a Reaction

Subscript	Meaning
r	any reaction specified
c	complete combustion
f	formation
d	decomposition
sol	solution
dil	dilution

For example, $\Delta_c H$ denotes the change in enthalpy (ΔH) of combustion (c) for a specified combustion reaction.

The value for the enthalpy change in Sample Problem 11.1 is given as a negative value ($\Delta_c H = -3.69$ kJ) because the change is a decrease. If the surroundings gained energy (i.e., $Q > 0$), then, according to the law of conservation of energy, the chemical system must have decreased in energy (i.e., $\Delta_c H < 0$). Use this reasoning to add the sign after doing the calculation.

Sample Problem 11.1 illustrates how we can infer the size of an enthalpy change from the size of a thermal energy change. Even though we cannot measure the value of an enthalpy change directly, we can determine it indirectly.

Chemists call a chemical change that produces energy an **exothermic reaction**. The enthalpy change (ΔH) is negative. A reaction that absorbs energy is called an **endothermic reaction**. The enthalpy change (ΔH) is positive.

Learning Tip

Remember **En**dothermic and **ex**othermic by associating the first two letters of **en**ter and **ex**it. Just remember that the energy is entering or exiting the chemical system, not the surroundings.

COMMUNICATION example 2

When 50 mL of 1.0 mol/L hydrochloric acid is neutralized completely by 75 mL of 1.0 mol/L sodium hydroxide in a polystyrene cup calorimeter, the temperature of the total solution changes from 20.2 °C to 25.6 °C. Determine the enthalpy change that occurs in the chemical system.

Solution

$$\begin{aligned}\Delta_r H &= Q \\ &= mc\Delta t \\ &= (50 + 75)\text{ g} \times \frac{4.19\text{ J}}{\text{g}\cdot^\circ\text{C}} \times (25.6 - 20.2)\,^\circ\text{C} \\ &= 2.83\text{ kJ}\end{aligned}$$

Based upon the evidence available, the enthalpy change for the neutralization of hydrochloric acid in this context is recorded as −2.83 kJ.

Learning Tip

There are several methods for solving these conservation of energy problems. The method shown here uses absolute values (with no positive or negative values in the equation). As a result, you have to provide the positive or negative sign to the final answer for ΔH.

WEB Activity

Case Study—Thermal Insulation

Living in Alberta in the winter involves dressing for the weather. Investigate what technological "advances" have been made by chemists creating synthetic materials. Compare these modern materials with traditional materials, such as animal skins, used for footwear and warm clothing.

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Molar Enthalpies and Calorimetry

Enthalpy of reaction (change in enthalpy) refers to the energy change for a whole chemical system when reactants change to products. At this stage in your study of thermochemistry you are looking only at energy change as a result of a chemical change in the chemical system. (You are not, for example, investigating energy changes due to phase changes, nuclear changes, or the dissolving of ionic compounds.) For all chemical changes a stated enthalpy change is for the total quantity of chemical involved in the reaction as specified by the reaction equation coefficients. Chemists sometimes want to know the enthalpy change per unit chemical amount (per mole) because they want to relate the enthalpy change to a chemical reaction equation that is balanced in moles.

Theoretically, **molar enthalpy of reaction** is the enthalpy change in a chemical system per unit chemical amount (per mole) of a specified chemical undergoing change in the system at constant pressure. Although enthalpy changes can be for any amount or mass of chemical undergoing a chemical change, molar enthalpy is the energy lost or gained expressed in units of, for example, kilojoules per mole (kJ/mol). The IUPAC symbol for molar enthalpy for a specified reaction is $\Delta_r H_m$.

Table 3 Quantities and Molar Quantities

Quantity (unit)	Molar quantity (unit)
mass, m (g)	molar mass, M (g/mol)
volume, V (L)	molar volume, V_m (L/mol)
enthalpy change, $\Delta_r H$ (kJ)	molar enthalpy, $\Delta_r H_m$ (kJ/mol)

In order to calculate enthalpy change from their respective molar quantities (**Table 3**), we have the following equation:

enthalpy change of reaction — chemical amount — molar enthalpy of reaction

$$\Delta_r H = n\,\Delta_r H_m$$

for example, $\Delta_c H = n\,\Delta_c H_m$ (c for combustion)

The latter equation is read as "the enthalpy of combustion is equal to the chemical amount of the fuel times the molar enthalpy of combustion of the fuel."

Molar quantities are often referenced (looked up in a table) or memorized: atomic molar masses are obtained from a periodic table; the molar volume of any gas at STP is memorized or referenced as 22.4 L/mol. We can use these molar quantities to predict, for example, the enthalpy of combustion ($\Delta_c H$) of an organic compound (as in Sample Problem 11.2).

DID YOU KNOW?

Measuring Enthalpy Change

Chemists and chemical engineers object to talking about *measuring* a change in enthalpy. They agree that we can measure a change in temperature (Δt) with a thermometer. Then we can *calculate* the change in thermal energy from $Q = mc\Delta t$. However, we must then use the law of conservation of energy to *infer* the molar enthalpy from the calculated thermal energy. In doing so, we *assume* that the change in enthalpy of the chemicals involved in a reaction is equal to the change in thermal energy of the surroundings.

Learning Tip

Note the logically consistent pattern for calculating quantities from their chemical amount (n) and molar quantity (Table 3). For example,

$$m = n\,M$$
$$= \text{mol} \times \text{g/mol}$$

$$V = n\,V_m$$
$$= \text{mol} \times \text{L/mol}$$

$$\Delta_r H = n\,\Delta_r H_m$$
$$= \text{mol} \times \text{kJ/mol}$$

We can see that the equations are correct because the units work out.

SAMPLE problem 11.2

Predict the change in enthalpy due to the combustion of 10.0 g of propane used in a camp stove (**Figure 6**).

First, you have to determine the change in enthalpy for the combustion, $\Delta_c H$, using the equation

$$\Delta_c H = n\,\Delta_c H_m$$

To calculate $\Delta_c H$ you need to know the chemical amount (n) and the molar enthalpy of combustion ($\Delta_c H_m$). You can calculate the chemical amount from the mass and molar mass. The molar enthalpy of combustion of propane (producing water vapour) is $-2\,043.9$ kJ/mol.

$$\Delta_c H = n\,\Delta_c H_m$$
$$= 10.0\text{ g} \times \frac{1\text{ mol}}{44.11\text{ g}} \times \frac{-2\,043.9\text{ kJ}}{1\text{ mol}}$$
$$= -463\text{ kJ}$$

Based on the concept of molar enthalpy, the change in enthalpy due to the combustion of 10.0 g of propane is -463 kJ.

Figure 6
Propane is a camp stove fuel. It can also be used for barbecues, lanterns, and recreation vehicle refrigerators.

COMMUNICATION example 3

Predict the enthalpy change due to the combustion of 10.0 g of butane in a camp heater. The molar enthalpy of combustion to produce water vapour is $-2\,657.3$ kJ/mol.

Solution

$$\Delta_c H = 10.0\text{ g} \times \frac{1\text{ mol}}{58.14\text{ g}} \times \frac{-2\,657.3\text{ kJ}}{1\text{ mol}}$$
$$= -457\text{ kJ}$$

Based upon the concept of molar enthalpy, the enthalpy change for the combustion of 10.0 g of butane is recorded as -457 kJ.

The molar enthalpy of reaction can be calculated from or used in calorimetry investigations. Again, the law of conservation of energy is used to produce an equality. The change in enthalpy (ΔH) of the chemical system is equal to the change in thermal energy (Q) of the calorimeter.

$$\Delta H = Q$$
$$n\,\Delta_c H_m = mc\Delta t$$

From this equation, any one of the five variables can be determined as an unknown. See Communication Example 4. The IUPAC symbols used in the calorimetry equation need to be interpreted carefully.

$$n\,\Delta_c H_m = mc\Delta t$$

- n: chemical amount of fuel
- $\Delta_c H_m$: molar enthalpy of combustion of fuel
- m: mass of water
- c: specific heat capacity of water
- Δt: temperature change of water

Learning Tip

The law of conservation of energy, as it relates to calorimetry, can be communicated as

$$\Delta H = Q$$
$$n\Delta_c H_m = mc\Delta t$$

However, the law of conservation of energy can be used without equating the two equations. Rather, we can first calculate the change in thermal energy, Q.

$$Q = mc\Delta t$$

then use the Q value as the $\Delta_r H$ value to calculate either the chemical amount or the molar enthalpy of reaction (as required); i.e.,

$$n = \frac{\Delta_c H}{\Delta_c H_m}$$

or

$$\Delta_c H_m = \frac{\Delta H}{n}$$

Your teacher may indicate whether you should write out these last two equations, or just substitute into them.

Learning Tip

There are several methods for solving these conservation of energy problems. The method shown in this textbook uses absolute values (with no positive or negative values in the equation). As a result, you have to provide the positive or negative sign to the final answer for ΔH. Negative values indicate exothermic changes and positive values indicate endothermic changes.

COMMUNICATION example 4

Ethanol is often added to gasoline as a renewable component that reduces harmful emissions. The mixture is known as gasohol. In a research laboratory, the combustion of 3.50 g of ethanol in a sophisticated calorimeter causes the temperature of 3.63 L of water to increase from 19.88 °C to 26.18 °C. Use this evidence to determine the molar enthalpy of combustion of ethanol.

Solution

$$Q = mc\Delta t$$

$$= 3.63 \text{ kg} \times \frac{4.19 \text{ J}}{\text{g}\cdot\text{°C}} \times (26.18 - 19.88) \text{ °C}$$

$$= 95.8 \text{ kJ}$$

$$\Delta_c H_{m\ C_2H_5OH} = \frac{95.8 \text{ kJ}}{3.50 \text{ g} \times \frac{1 \text{ mol}}{46.08 \text{ g}}}$$

$$= 1.26 \text{ MJ/mol}$$

Based upon the evidence gathered, the molar enthalpy of combustion of ethanol is recorded as -1.26 MJ/mol.

EXTENSION

Stronger Hurricanes

The theory of global warming is now accepted by a majority of scientists. But global warming doesn't just make our winters milder and our summers hotter. Find out how global warming might be completely changing weather patterns around the world.

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Practice

9. Write the symbols for the following terms.
(a) enthalpy of formation
(b) enthalpy of decomposition
(c) molar enthalpy of formation
(d) molar enthalpy of decomposition

10. In your own words, describe the similarities and differences between the enthalpy change and molar enthalpy symbols and meanings.

11. What is the significance of a positive or negative sign for an enthalpy change or molar enthalpy value?

12. Predict the enthalpy change for the combustion of every 100 g of methane in a natural gas water heater. ($\Delta_c H_m$ (CH_4) = -802.5 kJ/mol (to produce water vapour))

13. Methanol is one type of fuel that is used in fondue heaters. In an experiment using a simple tin can calorimeter, 2.98 g of methanol was burned to raise the temperature of 650 g of water by 20.9 °C. Using this evidence, calculate the molar enthalpy of combustion of methanol (to produce water as a vapour).

LAB EXERCISE 11.A

Molar Enthalpy of Neutralization

The molar enthalpy of neutralization, $\Delta_n H_m$, for hydrochloric acid when it reacts with a base can be determined by calorimetry. When two aqueous solutions are mixed in a calorimeter, the final aqueous mixture is considered to be the water in the calorimeter when calculating Q.

Report Checklist

○ Purpose	○ Design	● Analysis
● Problem	● Materials	○ Evaluation
○ Hypothesis	○ Procedure	
○ Prediction	○ Evidence	

Purpose

The purpose of this exercise is to use calorimetry to determine the molar enthalpy of neutralization for hydrochloric acid in its reaction with aqueous ammonia.

Design

1.00 L of 0.100 mol/L hydrochloric acid and 1.10 L of 0.100 mol/L aqueous ammonia are combined in a calorimeter.

Evidence

initial temperature of each solution = 23.20 °C
final temperature of mixture = 23.82 °C
volume of mixture = 2.10 L

WEB Activity

Simulation—Calorimetry

The purpose of this activity is to use a calorimetry computer model to investigate the quantity of heat produced by combustion of various masses of different compounds in excess oxygen. Keep all variables constant, other than changing a reactant.

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INVESTIGATION 11.2 *Introduction*

Molar Enthalpy of Reaction

The accuracy (percent difference) of the value for molar enthalpy obtained in this investigation is used to evaluate the calorimeter and the assumptions made in the analysis, rather than to evaluate the prediction and its authority. The ultimate authority in this experiment is considered to be the reference value used in the prediction.

Purpose

The purpose of this investigation is to test the calorimeter design and calorimetry procedure by comparing experimental evidence with a widely accepted value for the molar enthalpy of a neutralization reaction.

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	● Materials	● Evaluation (2, 1, 3)
○ Hypothesis	● Procedure	
○ Prediction	● Evidence	

Problem

What is the molar enthalpy of neutralization for sodium hydroxide when 50 mL of aqueous 1.0 mol/L sodium hydroxide reacts with an excess quantity of 1.0 mol/L sulfuric acid?

Prediction

The molar enthalpy of neutralization for sodium hydroxide is −57 kJ/mol, as per *The CRC Handbook of Chemistry and Physics*.

To perform this investigation, turn to page 516.

SUMMARY

- Thermal energy—the kinetic energy of entities (e.g., molecules of the surroundings, around a chemical system). Thermal energy increases with temperature.
- Change in thermal energy (Q in kJ)—measured by a temperature change and calculated using the relationship (equation) $Q = mc\Delta t$
- Heat—thermal energy transferred between systems. Heat is not possessed by a system. Heat is energy in transition (flowing) between systems.
- Kinetic energy (E_k in kJ)—a form of energy related to the motion of a chemical entity; measured as an average kinetic energy by the temperature of a chemical system
- Chemical potential energy (E_p in kJ)—the energy present in the chemical bonds of a substance
- Enthalpy change ($\Delta_r H$ in kJ)—a change in the enthalpy of a chemical system under constant pressure; also loosely called heat of reaction
- Molar enthalpy ($\Delta_r H_m$ in kJ/mol)—the change in enthalpy expressed per mole of a substance undergoing a specified reaction
- Endothermic change—a change in chemical energy where energy/heat enters (is absorbed by) the chemical system under consideration; results in an increase in chemical potential energy (and enthalpy); the $\Delta_r H$ is a positive value
- Exothermic change—a change in chemical energy where energy/heat exits (is released by) the chemical system under consideration; results in a decrease in chemical potential energy (and enthalpy); the $\Delta_r H$ is a negative value

Section 11.2 Questions

1. Record whether each of these statements is true or false. Provide your reasoning.
   (a) Heat is possessed by a chemical system.
   (b) For an exothermic change, energy exits from the surroundings and enters the chemical system.
   (c) The change in enthalpy of a chemical system is measured in kilojoules per mole (kJ/mol).
   (d) The change in thermal energy of the surroundings can be calculated using $Q = mc\Delta t$.
   (e) Temperature is the measure of the average kinetic energy of the entities in a chemical system.

2. It is commonly assumed in calorimetry labs with polystyrene calorimeters that a negligible quantity of heat is absorbed or released by the solid calorimeter materials such as the cup, stirring rod, and thermometer. Use the empirical data in **Table 4** to evaluate this assumption.
   (a) For a temperature change of 5.0 °C, calculate the energy change of the water only.
   (b) For a temperature change of 5.0 °C, calculate the total energy change of the water, polystyrene cups, stirring rod, and thermometer.
   (c) Calculate the percent error introduced by using only the energy change of the water.
   (d) Evaluate the assumption of negligible heat transfer to the solid calorimeter materials.

Table 4 Typical Quantities for Materials in a Simple Calorimeter

Material	Specific heat capacity (J/(g·°C))	Mass (g)
water	4.19	100.00
polystyrene cups	0.30	3.58
glass stirring rod	0.84	9.45
thermometer	0.87	7.67

3. Ethane is a component of natural gas. If the molar enthalpy of combustion of ethane is −1.56 MJ/mol, calculate the quantity of thermal energy gained by the burning of
   (a) 5.0 mol of ethane
   (b) 40.0 g of ethane

4. Decane is one of the hundreds of compounds in gasoline. The molar enthalpy of combustion of decane ($C_{10}H_{22}$) is −6.78 MJ/mol. What mass of decane would have to be burned in order to raise the temperature of 500.0 mL of water from 20.0 °C to 55.0 °C?

5. During sunny days, chemicals can store solar energy in homes for later release. Certain hydrated salts dissolve endothermically in their water of hydration when heated to decomposition and release heat when they solidify. For example, Glauber's salt, $Na_2SO_4 \cdot 10\ H_2O_{(s)}$, solidifies at 32 °C, releasing 78.0 kJ/mol of salt. What is the enthalpy change for the solidification of 1.00 kg of Glauber's salt used to supply energy to a home?

6. In a laboratory investigation into the neutralization reaction

   $HNO_3(aq) + KOH(s) \rightarrow KNO_3(aq) + H_2O(l)$

   a researcher adds solid potassium hydroxide to nitric acid solution in a polystyrene calorimeter. Complete the Analysis section of the lab report, including a calculation of the molar enthalpy of neutralization of potassium hydroxide.

   Evidence
   mass KOH = 5.2 g
   volume of nitric acid solution = 200 mL
   t_1 = 21.0 °C t_2 = 28.1 °C

7. In a laboratory investigation into the reaction

   $Ba(NO_3)_2(s) + K_2SO_4(aq) \rightarrow BaSO_4(s) + 2\ KNO_3(aq)$

   a researcher adds a 261 g sample of barium nitrate to 2.0 L of potassium sulfate solution in a polystyrene calorimeter. Complete the Analysis section of the lab report, including a calculation of the molar enthalpy of reaction of barium nitrate.

   Evidence
   As the barium nitrate dissolves, a precipitate is immediately formed.
   t_1 = 26.0 °C t_2 = 29.1 °C

8. A student noticed that chewing fast-energy dextrose tablets made her mouth feel cold. Write the Purpose, Problem, Hypothesis, and Design for an investigation to find out whether there really is a temperature change.

9. Plan an investigation to compare the molar enthalpies of combustion for butane and propane. Include a Problem statement, a Design, and a list of materials. In your Design, identify major variables and how you will control them, as well as any necessary precautions.

10. Describe the features of a calorimeter that make it appropriate for its purpose. What drawbacks does it have?

11. The molar enthalpy of neutralization, $\Delta_n H_m$, for a reactant in an aqueous acid–base reaction can be determined by calorimetry. Write the Problem, Procedure, and Analysis for the lab report.

    Purpose
    The purpose of this exercise it to use calorimetry to determine the molar enthalpy of neutralization for nitrous acid or for aqueous potassium hydroxide.

    Design
    1.00 L of 0.400 mol/L nitrous acid and 1.00 L of 0.400 mol/L aqueous potassium hydroxide are combined in a calorimeter.

    Evidence
    inital temperature of solutions = 22.08 °C
    final temperature of mixture = 24.04 °C

12. Research the technological applications of propane and butane. Identify the chemical properties of each that make these appropriate technologies for each of these hydrocarbons.

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Communicating Enthalpy Changes 11.3

Scientists and engineers have a variety of methods for communicating the energy change for chemical reactions. It is not uncommon to find different communication systems in different contexts. Each of these systems has pros and cons, but understanding all of them provides a deeper understanding of the topic of reaction enthalpies.

Most information about energy changes comes from the experimental method of calorimetry (page 485). These studies yield molar enthalpies that can be communicated in at least four ways:

1. by stating the molar enthalpy of a specific reactant in a reaction
2. by stating the enthalpy change for a balanced reaction equation
3. by including an energy value as a term in a balanced reaction equation
4. by drawing a chemical potential energy diagram

All four of these methods of expressing energy changes are equivalent. The first three are closer to empirical descriptions, and the fourth method is a theoretical description. Each of these methods of communicating energy changes in chemical reactions is described in the following sub-sections.

Method 1: Molar Enthalpies of Reaction, $\Delta_r H_m$

The molar enthalpy of reaction for a substance is determined empirically from the measurement of the energy released or absorbed per unit chemical amount to or from the surroundings of the chemical system at constant pressure. The energy measurements usually involve calorimetry. To communicate a molar enthalpy, both the substance and the reaction must be specified. The substance is conveniently specified by its chemical formula. Some chemical reactions are well known and specific enough to be identified by name only. For instance, reference books often list standard molar enthalpies of formation ($\Delta_f H_m°$) and combustion ($\Delta_c H_m°$).

Standard enthalpy values are expressed with a ° superscript, as in $\Delta_f H_m°$. The reactants and products are in their standard state: a pressure of 100 kPa, an aqueous concentration of 1 mol/L, and liquids and solids in their pure state. For liquid and solid elements the pure state must be at 25 °C. Liquid and solid compounds must only have the same initial and final temperature (most often 25 °C). For our purposes we often refer to SATP as the standard conditions for both elements and compounds, unless special conditions exist for the liquid and solid compounds.

For some well-known reactions, such as formation and combustion, no chemical equation is necessary, since formation and combustion refer to specific, unambiguous chemical reactions indicated by the subscript letter after the delta symbol. For example, the molar enthalpy of formation for methanol at standard conditions is communicated internationally as

$$\underset{CH_3OH}{\Delta_f H_m°} = -239.2 \text{ kJ/mol}$$

This means that 239.2 kJ of energy is released to the surroundings when 1 mol of methanol is formed from its elements when they are in their standard states at SATP. The following chemical equation communicates the formation reaction assumed to occur:

$$C(s) + 2\,H_2(g) + \tfrac{1}{2}\,O_2(g) \rightarrow CH_3OH(l)$$

DID YOU KNOW ?

Communication

Communication is a very important part of any discipline. Here are some words of wisdom from famous scientists to their students.

No amount of experiments can ever prove me right; a single experiment may at any time prove me wrong.

Imagination is more important than knowledge.

It is the theory which decides what we can observe.

Albert Einstein

Truth in science can be defined as the working hypothesis best suited to open the way to the next better one.

Konrad Lorentz

Discovery is seeing what everybody else has seen, and thinking what nobody else has thought.

Albert Szent-Györgyi

The foolish reject what they see, not what they think; the wise reject what they think, not what they see.

Huang Po

Science is a way to teach how something gets to be known, what is not known, to what extent things are known (for nothing is known absolutely), how to handle doubt and uncertainty, what the rules of evidence are, how to think about things so that judgments can be made, and how to distinguish truth from fraud, and from show.

Richard Feynman

Where observation is concerned, chance favours only the prepared mind.

Louis Pasteur

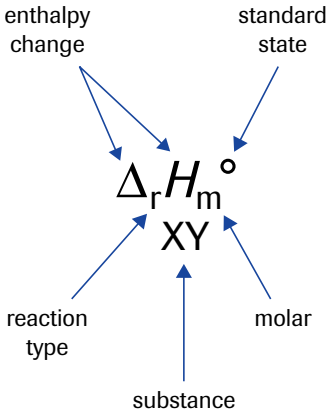


Figure 1
An explanation of the parts of an enthalpy symbol

Figure 2
Methanol burns more completely than gasoline and produces lower levels of some pollutants. The technology of methanol-burning vehicles was originally developed for racing cars because methanol burns faster than gasoline. However, its energy content is lower so that it takes twice as much methanol as gasoline to drive a given distance.

DID YOU KNOW ?

Renewable Energy Sources

Partly in reaction to the health hazards of some components of fossil fuels, scientists and engineers around the world are searching for alternative energy sources. Energy sources with low health, environmental, and resource impact include hydroelectric, wind, solar, biomass, and geothermal energy, along with small-scale hydro and tidal. Research by Statistics Canada indicates that, at the beginning of the 21st century, wind is the fastest-growing form of renewable energy in Canada.

As you know, we can find molar enthalpies from measurements involving systems at various conditions of temperature and pressure. If, however, chemists determine a molar enthalpy when the initial and final conditions of the chemical system are standard conditions, they call it a **standard molar enthalpy of reaction**. The symbol $\Delta_r H_m°$ distinguishes standard molar enthalpies from molar enthalpies, $\Delta_r H_m$ (**Figure 1**). Standard molar enthalpies allow chemists to create tables to compare enthalpy values and to increase the precision of frequently used values by careful experimentation. For an exothermic reaction, the standard molar enthalpy is calculated by taking into account all the energy required to change the reaction system from standard conditions in order to initiate the reaction *and* all the energy released following the reaction, as the products are cooled to standard conditions. For example, the standard molar enthalpy of combustion of methanol is

$$\underset{CH_3OH}{\Delta_c H_m°} = -725.9 \text{ kJ/mol}$$

This means that the complete combustion of 1 mol of methanol (**Figure 2**) releases 725.9 kJ of energy according to the following balanced equation:

$$CH_3OH(l) + \tfrac{3}{2} O_2(g) \rightarrow CO_2(g) + 2 H_2O(l)$$

For a *standard* value, the initial and final conditions of the chemical system must be standard state. In this case, the carbon dioxide and liquid water are produced at a high temperature. They would be allowed to cool in a special calorimeter to standard conditions before experimenters take the final measurement in order to calculate the energy produced.

If a chemical reaction is not well known or if the equation for the reaction is not obvious, then the chemical equation must be stated along with the molar enthalpy. For example, methanol is produced industrially by the high-pressure reaction of carbon monoxide and hydrogen gases.

$$CO(g) + 2 H_2(g) \rightarrow CH_3OH(l)$$

Chemists have determined that, in this rection, the standard molar enthalpy of reaction for methanol, $\Delta_r H_m°$, is -128 kJ/mol. Note that this is not a formation reaction since the reactants are not all elements. Note also that you must always state which substance you are considering for this method of communication.

$$\underset{CH_3OH}{\Delta_r H_m°} = -128 \text{ kJ/mol}$$

Method 2: Enthalpy Changes, $\Delta_r H$

A second method for communicating an energy change is to write an enthalpy change ($\Delta_r H$) beside the chemical equations. For example,

$$CO(g) + 2 H_2(g) \rightarrow CH_3OH(l) \qquad \Delta_r H° = -128 \text{ kJ}$$

Molar enthalpies of reaction can be used to calculate the enthalpy change during a chemical reaction; a molar enthalpy and a balanced chemical equation are required for the calculation. The enthalpy change is calculated using the empirical definition,

$$\Delta_r H° = n \Delta_r H_m°$$

where n is the chemical amount (in moles) of the substance whose molar enthalpy is known. *The enthalpy change for the equation as written equals the chemical amount (from the coefficient in the equation) times the molar enthalpy of reaction (for a specific chemical).* See Sample Problem 11.3.

SAMPLE problem 11.3

Sulfur dioxide and oxygen react to form sulfur trioxide (**Figure 3**). The standard molar enthalpy of combustion of sulfur dioxide, in this reaction, is −98.9 kJ/mol. What is the enthalpy change for this reaction? First write the balanced chemical equation.

$$2\,SO_2(g) + O_2(g) \rightarrow 2\,SO_3(g)$$

Then obtain the chemical amount of sulfur dioxide from its coefficient in the balanced equation and use $\Delta_c H° = n\,\Delta_c H_m°$.

$$\Delta_c H° = n\,\Delta_c H_m°$$

$$= 2\text{ mol} \times \frac{-98.9\text{ kJ}}{1\text{ mol}} = -197.8\text{ kJ}$$

Report the enthalpy change for the reaction by writing it next to the balanced equation, as follows:

$$2\,SO_2(g) + O_2(g) \rightarrow 2\,SO_3(g) \qquad \Delta_c H° = -197.8\text{ kJ}$$

Note that the *2* in *2 mol* is an exact number; it does not affect the certainty of the answer.

Figure 3
Most sulfuric acid is produced in plants like this by the contact process, which includes two exothermic combustion reactions. Sulfur reacts with oxygen, forming sulfur dioxide; sulfur dioxide, in contact with a catalyst, reacts with oxygen, forming sulfur trioxide.

Note that the enthalpy change is not, in this method, a molar value, so does not require the "m" subscript and is not in kJ/mol. Also, it is not necessary to state which substance is under consideration because the quantity refers to the reaction equation *as written*.

The enthalpy change depends on the actual chemical amount of reactants and products in the chemical reaction. Therefore, if the balanced equation for the reaction is written differently, the enthalpy change should be reported differently. For example,

$$SO_2(g) + \tfrac{1}{2}\,O_2(g) \rightarrow SO_3(g) \qquad \Delta_c H° = -98.9\text{ kJ}$$

$$2\,SO_2(g) + O_2(g) \rightarrow 2\,SO_3(g) \qquad \Delta_c H° = -197.8\text{ kJ}$$

Both chemical equations agree with the empirically determined *molar* enthalpy of combustion for sulfur dioxide in this reaction.

$$\Delta_c H_{m,\,SO_2}° = \frac{-197.8\text{ kJ}}{2\text{ mol}} = -98.9\text{ kJ/mol}$$

Unlike molar enthalpies of formation or combustion, the enthalpy changes for most reactions must be accompanied by a balanced chemical equation so that we know what reaction is being described.

Learning Tip

If a chemical equation is altered by multiplying or dividing by some factor, then the $\Delta_r H$ is altered in exactly the same way.

COMMUNICATION example 1

Wild natural gas wells are sometimes lit on fire to eliminate the very toxic hydrogen sulfide gas. The standard molar enthalpy of combustion of hydrogen sulfide is −518.0 kJ/mol. Express this value as a standard enthalpy change for the following reaction equation:

$$2\,H_2S(g) + 3\,O_2(g) \rightarrow 2\,H_2O(g) + 2\,SO_2(g) \qquad \Delta_c H° = ?$$

Solution

$$\Delta_c H = 2\text{ mol} \times -518.0\text{ kJ/mol}$$

$$= -1\,036.0\text{ kJ}.$$

$$2\,H_2S(g) + 3\,O_2(g) \rightarrow 2\,H_2O(g) + 2\,SO_2(g) \qquad \Delta_c H° = -1\,036.0\text{ kJ}$$

Learning Tip

Fractions are convenient in many thermochemical equations but are not recommended by IUPAC. Note that use of fractions applies to fractions of moles of substances (e.g., $\frac{3}{2}$ mol represents 1.5 mol) rather than fractions of actual molecules.

Oxygen is often the reactant given a fractional coefficient in combustion equations. This is because it occurs as a diatomic molecule and the total number of oxygen atoms in the products is often an odd number.

Figure 4
Combustion reactions are the most familiar exothermic reactions.

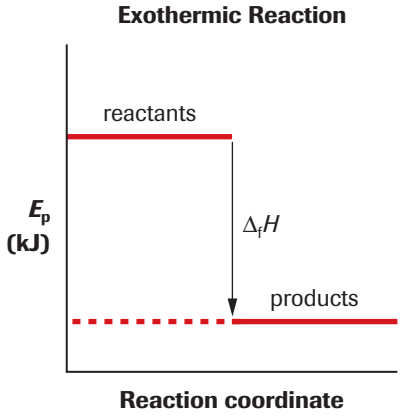


Figure 5
During an exothermic reaction, the enthalpy of the system decreases and heat flows into the surroundings. We observe a temperature increase in the surroundings.

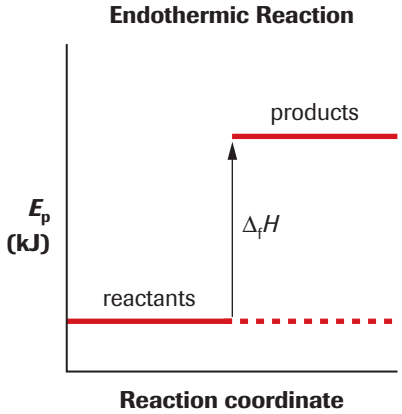


Figure 6
During an endothermic reaction, heat flows from the surroundings into the chemical system. We observe a temperature decrease in the surroundings.

Method 3: Energy Terms in Balanced Equations

Another way to report the enthalpy change in a chemical reaction is to include it as a term in a balanced equation. If a reaction is endothermic, it requires a certain quantity of additional energy for the reactants to continuously react. This energy (like the reactants) is transformed as the reaction progresses and is listed along with the reactants. For example,

$$H_2O(l) + 285.8\ kJ \rightarrow H_2(g) + \tfrac{1}{2} O_2(g)$$

If a reaction is exothermic, energy is released as the reaction proceeds (**Figure 4**) and is listed along with the products. For example,

$$Mg(s) + \tfrac{1}{2} O_2(g) \rightarrow MgO(s) + 601.6\ kJ$$

In order to specify the initial and final conditions for measuring the enthalpy change of the reaction, the temperature and pressure may be specified at the end of the equation.

$$Mg(s) + \tfrac{1}{2} O_2(g) \rightarrow MgO(s) + 601.6\ kJ \qquad \text{(under standard conditions)}$$

COMMUNICATION example 2

Ethane is cracked into ethene in world-scale quantities in Alberta. Communicate the enthalpy of reaction as a term in the equation representing the cracking reaction.

$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g) \qquad \Delta_{cr}H = +136.4\ kJ$

Solution

$C_2H_6(g) + 136.4\ kJ \rightarrow C_2H_4(g) + H_2(g)$

Note that, in Communication Example 2, the enthalpy change has a positive value. This tells us that the reaction is endothermic, so the energy term is written with the reactants.

Method 4: Chemical Potential Energy Diagrams

Chemists use the law of conservation of energy to describe what happens during a chemical reaction. They explain their observations theoretically: observed energy changes are due to changes in chemical potential energy that occur during a reaction. This energy is a stored form of energy that is related to the relative positions of particles and the strengths of the bonds between them. As bonds break and re-form and the positions of atoms are altered, changes in potential energy occur. Evidence of a change in enthalpy of a chemical system is provided by a temperature change of its surroundings.

A **chemical potential energy diagram** shows the potential energy of both the reactants and the products of a chemical reaction (**Figures 5** and **6**). The difference between the initial and final energies in a chemical potential energy diagram is the enthalpy change, obtained from calorimetry by measuring the temperature change of the calorimeter. A temperature change is caused by a flow of heat into or out of the chemical system.

The vertical axis on the diagram represents the potential energy of the system, E_p. The reactants are written on the left and the products on the right, and the horizontal axis is called the reaction coordinate or reaction progress. In an exothermic change (**Figure 5**), the products have less potential energy than the reactants: energy is released to the surroundings as the products form. In an endothermic change (**Figure 6**), the

products have more potential energy than the reactants: energy is absorbed from the surroundings. Neither of the axes is numbered; only the numerical change in potential energy (enthalpy change, $\Delta_r H$) of the system is shown in the diagrams.

COMMUNICATION example 3

Communicate the following enthalpies of reaction as chemical potential energy diagrams.
(a) The burning of magnesium to produce a very bright emergency flare.

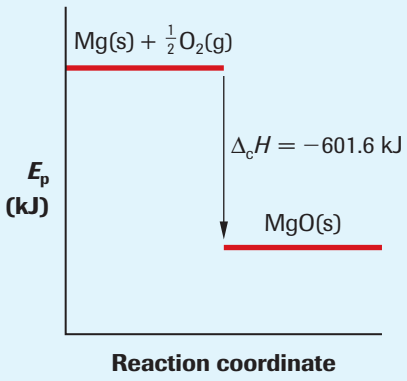
$$Mg(s) + \tfrac{1}{2}O_2(g) \rightarrow MgO(s) \qquad \Delta_c H = -601.6 \text{ kJ}$$

(b) The decomposition of water by electrical energy from a solar cell.

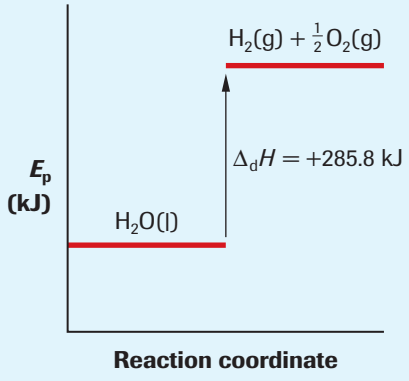
$$\Delta_d H_m = +285.8 \text{ kJ/mol}$$

Solution

(a) **Combustion of Magnesium**

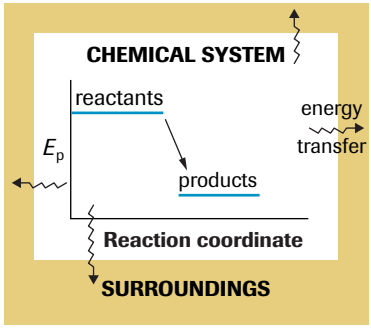


(b) **Decdecomposition of Water**



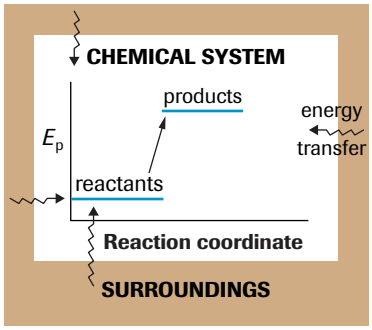
Each of the four methods of communicating the molar enthalpy or change in enthalpy of a chemical reaction has advantages and disadvantages. To best understand energy changes in chemical reactions, you should learn all four methods. Communication Example 4, on the next page, illustrates these methods for an exothermic and an endothermic reaction (**Figure 7**).

(a) Exothermic Reaction



Surroundings are warmed as chemical system releases energy.

(b) Endothermic Reaction



Surroundings are cooled as chemical system absorbs energy.

Figure 7
According to the law of conservation of energy, energy is neither created nor destroyed, only transformed. The diagrams above convey the notion that energy exits from an exothermic reaction system and enters into an endothermic reaction system.

SUMMARY Four Ways of Communicating Energy Changes

	Exothermic Changes	Endothermic Changes
1. Molar Enthalpy	$\Delta_r H_m < 0$	$\Delta_r H_m > 0$
2. Enthalpy Change	reactants $\rightarrow$ products; $\Delta_r H < 0$	reactants $\rightarrow$ products; $\Delta_r H > 0$
3. Term in a Balanced Equation	reactants $\rightarrow$ products + energy	reactants + energy $\rightarrow$ products
4. Chemical Potential Energy Diagram	E_p (reactants) > E_p (products)	E_p (reactants) < E_p (products)

Learning Tip

Remember that the unit of $\Delta_r H$ is the kJ and the unit of $\Delta_r H_m$ is the kJ/mol, and that when multiplying by an exact number, you use the precision rule. For example, 2×40.65 kJ $=$ 81.30 kJ, retaining the same precision as the measured value.

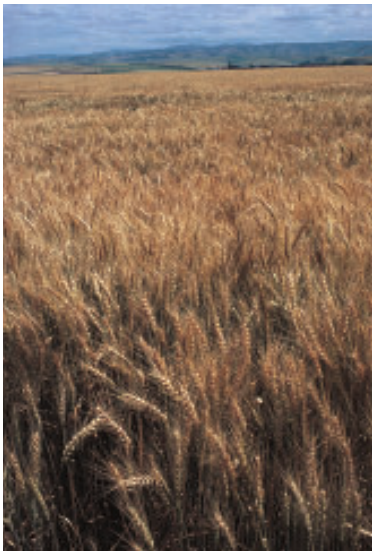


Figure 8
Energy can be captured by and from grain. The energy may be obtained from the grain in the form of, for example, foodstuffs, ethanol, and/or biodiesel. Grain grows well during the long summer days in Alberta, when the photosynthesis process exceeds the cellular respiration process. A grain field is an example of an open system, in which mass and energy flow freely into and out of the system.

COMMUNICATION example 4

Energy is transformed in cellular respiration and in photosynthesis (**Figure 8**). Cellular respiration, a series of exothermic reactions, is the breakdown of foodstuffs, such as glucose, that takes place within cells. Photosynthesis, a series of endothermic reactions, is the process by which green plants use light energy to make glucose from carbon dioxide and water.

Express the standard enthalpy changes for cellular respiration and for photosynthesis by using the four different methods of communication.

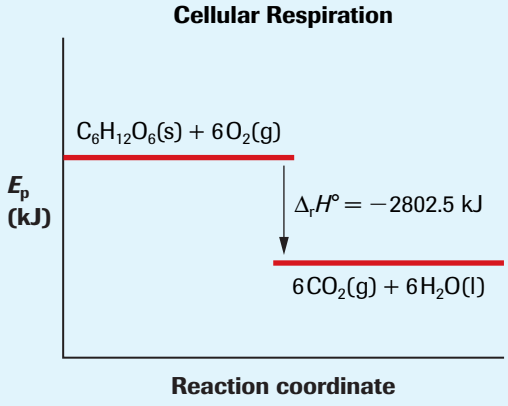
(a) One mole of glucose is consumed, during cellular respiration, to release 2802.5 kJ of energy.

(b) Glucose is produced during photosynthesis.

Solution

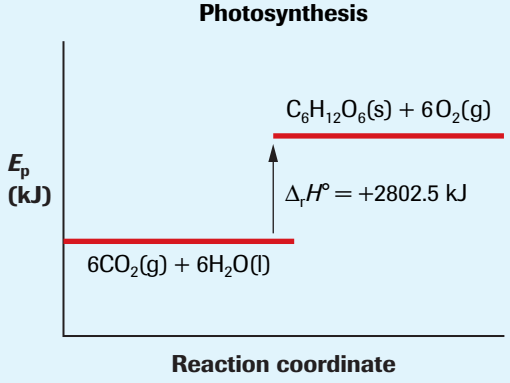
(a)

1. The standard molar enthalpy for cellular respiration of glucose:
   $\Delta_r H_m^\circ = -2802.5$ kJ/mol
   $C_6H_{12}O_6$
2. $C_6H_{12}O_6(g) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) \quad \Delta_r H^\circ = -2802.5$ kJ
3. $C_6H_{12}O_6(s) + 6O_2(g) \rightarrow 6CO_2(g) + 6H_2O(l) + 2802.5$ kJ
4. Potential energy diagram for cellular respiration:



(b)

1. The standard molar enthalpy for photosynthesis of glucose:
   $\Delta_r H^\circ_m = +2802.5$ kJ/mol
   $C_6H_{12}O_6$
2. $6CO_2(g) + 6H_2O(l) \rightarrow C_6H_{12}O_6(s) + 6O_2(g) \quad \Delta_r H^\circ = +2802.5$ kJ
3. $6CO_2(g) + 6H_2O(l) + 2802.5$ kJ $\rightarrow C_6H_{12}O_6(s) + 6O_2(g)$
4. Potential energy diagram for photosynthesis:



Section 11.3 Questions

1. Translate into words all parts of the following symbols:
(a) $\Delta_c H_m°$
CH_4
(b) $n\,\Delta_f H_m°$
C_3H_8
(c) Q
water

2. Communicate the standard enthalpy change by using the four methods described in this section for each of the following chemical reactions. Assume standard conditions for the measurements of initial and final states.
(a) The formation of acetylene (ethyne, C_2H_2) fuel from solid carbon and gaseous hydrogen ($\Delta_f H_m° = +227.4$ kJ/mol acetylene)
(b) The simple decomposition of aluminium oxide powder ($\Delta_{sd} H_m° = +1675.7$ kJ/mol aluminium oxide)
(c) The complete combustion of pure carbon fuel ($\Delta_c H_m° = -393.5$ kJ/mol CO_2)

3. For each of the following balanced chemical equations and standard enthalpy changes, write the symbol and calculate the standard molar enthalpy of combustion for the substance that reacts with oxygen.
(a) $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$ $\quad \Delta_c H° = -483.6$ kJ
(b) $4\,NH_3(g) + 7\,O_2(g) \rightarrow 4\,NO_2(g) + 6\,H_2O(g) + 1272.1$ kJ
(c) $2\,N_2(g) + O_2(g) + 163.2$ kJ $\rightarrow 2\,N_2O(g)$
(d) $3\,Fe(s) + 2\,O_2(g) \rightarrow Fe_3O_4(s)$ $\quad \Delta_c H° = -1118.4$ kJ

4. The neutralization of a strong acid and a strong base is an exothermic process.

$$H_2SO_4(aq) + 2\,NaOH(aq) \rightarrow Na_2SO_4(aq) + 2\,H_2O(l) + 114\text{ kJ}$$

(a) What is the enthalpy change for this reaction?
(b) Write this chemical equation, using the $\Delta_r H$ notation not under standard conditions.
(c) Calculate the molar enthalpy of neutralization for sulfuric acid.
(d) Calculate the molar enthalpy of neutralization for sodium hydroxide.

5. Translate the empirical molar enthalpies given below into a balanced chemical equation, including the standard enthalpy change; for example,

$$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)\ \Delta_c H° = -802.5\text{ kJ}$$

(a) The standard molar enthalpy of combustion for methanol to produce water vapour is -725.9 kJ/mol.
(b) The standard molar enthalpy of formation for liquid carbon disulfide is 89.0 kJ/mol.
(c) The standard molar enthalpy of roasting (combustion) for zinc sulfide is -441.3 kJ/mol.
(d) The standard molar enthalpy of simple decomposition, $\Delta_{sd} H_m°$, for iron(III) oxide is 824.2 kJ/mol.

6. For each of the following reactions, translate the given standard molar enthalpy into a balanced chemical equation using the $\Delta_r H°$ notation, and then rewrite the equation including the energy as a term in the equation. Assume water is produced as a vapour.
(a) Propane obtained from natural gas is used as a fuel in barbecues and vehicles (**Figure 9**).
$\Delta_c H_m° = -2\,043.9$ kJ/mol C_3H_8
(b) Nitrogen monoxide forms at the high temperatures inside an automobile engine. $\Delta_f H_m° = +91.3$ kJ/mol NO
(c) Some advocates of alternative fuels have suggested that cars could run on ethanol.
$\Delta_c H_m° = -1\,234.8$ kJ/mol C_2H_5OH

Figure 9
Propane-fuelled vehicles are not allowed to park in underground parking lots. Propane is denser than air, and a dangerous quantity of propane could accumulate in the event of a leak.

7. The standard molar enthalpy of combustion for hydrogen to produce liquid water is -285.8 kJ/mol. The standard molar enthalpy of decomposition for liquid water is $+285.8$ kJ/mol.
(a) Write both chemical equations using the $\Delta_r H°$ notation.
(b) How does the standard enthalpy change for the combustion of hydrogen compare with the standard enthalpy change for the simple decomposition of liquid water? Suggest a generalization to include all pairs of chemical equations that are the reverse of one another.

8. Combustion reactions are very useful to society, but also have some drawbacks. Create a list of possible drawbacks of our reliance on combustion reactions, and suggest some possible solutions.

11.4 Hess' Law

DID YOU KNOW ?

Germain H. Hess and Enthalpy

An amazing part of the history of Hess' law is that it was created before the law of conservation of energy was well formulated. Hess' law also preceded the use of the term *enthalpy,* which did not appear until nearly fifty years later. Hess is known as the founder of thermochemistry for his extensive empirical work with heat produced by neutralization and combustion reactions.

Calorimetry is the basis for most information about chemical energy changes. However, not every reaction of interest to scientists and engineers can be studied by means of a calorimetric experiment. For example, the rusting of iron is extremely slow and, therefore, results in temperature changes too small to be measured using a conventional calorimeter. It is similarly impossible to measure, with a calorimeter, the temperature change involved in the energy of formation of carbon monoxide because the combustion of carbon produces carbon dioxide and carbon monoxide simultaneously. Chemists have devised a number of methods to predict an enthalpy change for reactions that are inconvenient to study experimentally. All the methods are based on the law of conservation of energy and the experimentally established principle that *net changes in all properties of a system are independent of the way the system changes from the initial state to the final state.* A temperature change is an example of a property that satisfies this principle. A net (overall) temperature change ($t_f - t_i$) does not depend on whether the temperature changed slowly, quickly, or rose and fell several times between the initial temperature and the final temperature. This same principle applies to enthalpy changes. If several reactions occur in different ways but the initial reactants and final products are the same, the net enthalpy change is the same as long as the reactions have the same initial and final conditions.

Predicting $\Delta_r H$: Hess' Law

Based on experimental measurements and calculations of enthalpy changes, Swiss chemist G.H. Hess suggested in 1840 that the addition of chemical equations yields a net chemical equation whose enthalpy change is the sum of the individual enthalpy changes. This generalization has been tested in many experiments and is now accepted as the law of additivity of enthalpies of reaction, also known as **Hess' law**. Hess' law can be written as an equation using the uppercase Greek letter Σ (pronounced "sigma") to mean "the sum of."

$$\Delta_r H = \Delta_1 H + \Delta_2 H + \Delta_3 H + \ldots$$

or

$$\Delta_r H = \Sigma \Delta_r H$$

or, in standard conditions,

$$\Delta_r H^\circ = \Sigma \Delta_r H^\circ$$

Figure 1
In this illustration, bricks on a construction site are being moved from the ground up to the second floor, but there are two different paths that the bricks could follow. In one path, they could move from the ground up to the third floor, and then be carried down to the second floor. In the other path, they would be carried up to the second floor in a single step. In both cases, the overall change in position—one floor up—is the same.

An analogy for this concept is shown in **Figure 1**. The net vertical distance that the bricks rise is the same whether they go up in one stage or in two stages. The same principle applies to enthalpy changes: If a set of reactions occurs in different steps but the initial reactants and final products are the same, the overall enthalpy change is the same.

Hess' discovery allowed the determination of the enthalpy change of a reaction without direct calorimetry (see Sample Problem 11.4), using two rules for chemical equations and enthalpy changes that you already know:

- If a chemical equation is reversed, then the sign of $\Delta_r H$ changes.
- If the coefficients of a chemical equation are altered by multiplying or dividing by a constant factor, then the $\Delta_r H$ is altered by the same factor.

SAMPLE problem 11.4

Use Hess' law to determine the standard enthalpy change for the formation of carbon monoxide.

$$C(s) + \tfrac{1}{2}O_2(g) \rightarrow CO(g) \qquad \Delta_f H° = ?$$

This reaction cannot be studied calorimetrically since the combustion of carbon produces carbon dioxide as well as carbon monoxide. However, the standard enthalpy of complete combustion for carbon and for carbon monoxide can be calculated by calorimetric measurements, and the standard enthalpy of formation for carbon monoxide can be determined using Hess law as follows:

$$(1)\quad C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta_c H° = -393.5 \text{ kJ}$$

$$(2)\quad 2CO(g) + O_2(g) \rightarrow 2CO_2(g) \qquad \Delta_c H° = -566.0 \text{ kJ}$$

Rearrange these two equations, and then add them together to obtain the chemical equation for the formation of carbon monoxide. The first term in the formation equation for carbon monoxide is 1 mol of solid carbon. Therefore, leave equation (1) unaltered so that C(s) will appear on the reactant side when we add the equations. However, we want 1 mol of CO(g) to appear as a product, so reverse equation (2) and divide each of its terms (including the standard enthalpy change) by 2.

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta_1 H° = -393.5 \text{ kJ}$$

$$CO_2(g) \rightarrow CO(g) + \tfrac{1}{2}O_2(g) \qquad \Delta_2 H° = +283.0 \text{ kJ}$$

Note that the sign of the standard enthalpy change in equation (2) has changed, since the equation has been reversed. Now add the reactants, products, and standard enthalpy changes to get a net reaction equation. Note that $CO_2(g)$ can be cancelled because it appears on both sides of the net equation. Similarly, $\frac{1}{2}O_2(g)$ can be cancelled from each side of the equation, resulting in:

$$C(s) + O_2(g) \rightarrow CO_2(g) \qquad \Delta_1 H° = -393.5 \text{ kJ}$$

$$CO_2(g) \rightarrow CO(g) + \tfrac{1}{2}O_2(g) \qquad \Delta_2 H° = +283.0 \text{ kJ}$$

$$C(s) + \tfrac{1}{2}O_2(g) \rightarrow CO(g) \qquad \Delta_f H° = -110.5 \text{ kJ}$$

While manipulating equations (1) and (2), you should check the desired equation and plan ahead to ensure that the substances end up on the correct sides and in the correct chemical amounts.

Learning Tip

The process for using Hess' law is a combination of being systematic and using trial and error. The given equations must be manipulated (like a system of equations in mathematics) to get what you want.

In Sample Problem 11.4 (as the accompanying audio clip explains):

- leave equation (1) as is
- divide equation (2) by 2 and reverse it.

Do what needs to be done to the given equations so they add to get the equation you want.

DID YOU KNOW?

Carbon Monoxide Poisoning

Carbon monoxide is known as the silent killer. Incomplete combustion can occur with deadly effects in the following circumstances:

- using charcoal briquettes in a home fireplace, recreational vehicle, or tent
- operating a car with a leaky exhaust system
- operating a fireplace in a home with inadequate ventilation (causing furnace exhaust to reverse into the home).

Carbon monoxide is odourless. Know how to prevent exposure to it, and install a carbon monoxide detector in your home or recreation vehicle.

Sketching a chemical potential energy diagram (**Figure 2**) might help you ensure that you have made the appropriate additions and subtractions.

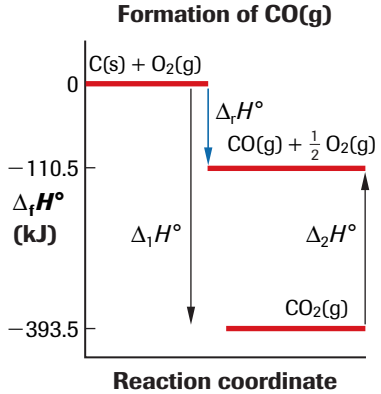


Figure 2
Carbon and oxygen react, forming carbon dioxide ($\Delta_1 H° = -393.5$ kJ), which theoretically could react to form carbon monoxide and oxygen ($\Delta_2 H° = +283.0$ kJ). The net standard enthalpy change of the two-step reaction is -393.5 kJ + 283.0 kH = -110.5 kJ.
Note that the label on the *y*-axis must be "standard enthalpy of formation, $\Delta_f H°$."

Figure 3
The molar enthalpy of combustion for butane can be determined calorimetrically. The molar enthalpy of formation, however, is best determined with Hess' law.

COMMUNICATION example

In an experiment to find the standard molar enthalpy of formation of butane (from its elements), the following values were determined by calorimetry:

(1) $C_4H_{10}(g) + \frac{13}{2}O_2(g) \rightarrow 4\,CO_2(g) + 5\,H_2O(g)$ $\quad \Delta_c H° = -2657.4$ kJ

(2) $C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta_f H° = -393.5$ kJ

(3) $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$ $\quad \Delta_f H° = -483.6$ kJ

What is the standard molar enthalpy of formation of butane (**Figure 3**)?

Solution

Net $\quad 4\,C(s) + 5\,H_2(g) \rightarrow C_4H_{10}(g)$ $\quad \Delta_f H° = ?$

(1) $4\,CO_2(g) + 5\,H_2O(g) \rightarrow C_4H_{10}(g) + \frac{13}{2}O_2(g)$ $\quad \Delta_1 H° = +2657.4$ kJ

(2) $\quad 4\,C(s) + 4\,O_2(g) \rightarrow 4\,CO_2(g)$ $\quad \Delta_2 H° = -1574.0$ kJ

(3) $\quad 5\,H_2(g) + \frac{5}{2}O_2(g) \rightarrow 5\,H_2O(g)$ $\quad \Delta_3 H° = -1209.0$ kJ

Net $\quad 4\,C(s) + 5\,H_2(g) \rightarrow C_4H_{10}(g)$ $\quad \Delta_f H° = -125.6$ kJ

$$\Delta_f H_m° \text{ (}C_4H_{10}\text{)} = \frac{-125.6\text{ kJ}}{1\text{ mol}}$$

$$= -125.6\text{ kJ/mol}$$

According to Hess' law and the data provided, the standard molar enthalpy of formation of butane is −125.6 kJ/mol.

Learning Tip

To obtain the formation equation and its enthalpy change in this Communication Example:

- reverse equation (1) and change the sign of the ΔH
- multiply equation (2) and its ΔH by 4
- multiply equation (3) and its ΔH by $\frac{5}{2}$

Figure 4
The reaction of powdered aluminium and iron(III) oxide is known as the "thermite" reaction and is very exothermic. Molten, white-hot iron is produced. In this apparatus, widely used in the past, the thermite reaction occurs in the upper chamber, and molten iron flows down into a mold between the ends of two rails, welding them together.

SUMMARY Enthalpy of Reaction and Hess' Law

To determine an enthalpy change of a reaction by using Hess' law, follow these steps:

1. Write the net reaction equation, if it is not given.
2. Manipulate the given equations so they will add to yield the net equation.
3. Multiply, divide, and/or reverse the sign of the enthalpy of reaction.
4. Cancel and add the remaining reactants and products to yield the net equation.
5. Add the component enthalpy changes to obtain the net enthalpy change.
6. Determine the molar enthalpy for a reactant or product, if required.

Practice

1. The standard enthalpy changes for the formation of aluminium oxide and iron(III) oxide (**Figure 4**) are

$2\,Al(s) + \frac{3}{2}O_2(g) \rightarrow Al_2O_3(s)$ $\quad \Delta_f H° = -1675.7$ kJ

$2\,Fe(s) + \frac{3}{2}O_2(g) \rightarrow Fe_2O_3(s)$ $\quad \Delta_f H° = -824.2$ kJ

Use Hess' law to predict the standard enthalpy change for the following reaction.

$Fe_2O_3(s) + 2\,Al(s) \rightarrow Al_2O_3(s) + 2\,Fe(s)$ $\quad \Delta_r H° = ?$

Use this information about coal gasification to answer questions 2 to 4.

Question 2:
Coal gasification converts coal into a combustible mixture of carbon monoxide and hydrogen, called *coal gas*, in a gasifier (**Figure 5**).

(1) $H_2O(g) + C(s) \rightarrow CO(g) + H_2(g)$ $\quad \Delta_r H° = ?$

Question 3:
The coal gas can be used as a fuel, for example, in a combustion turbine.

(2) $CO(g) + H_2(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ $\quad \Delta_c H° = ?$

Question 4:
As an alternative to combustion, coal gas can undergo a process called *methanation*.

(3) $3\,H_2(g) + CO(g) \rightarrow CH_4(g) + H_2O(g)$ $\quad \Delta_r H° = ?$

2. Calculate the standard enthalpy change for reaction (1)—the gasification of coal—from the following chemical equations and standard enthalpy changes.

$2C(s) + O_2(g) \rightarrow 2CO(g)$ $\quad \Delta_f H° = -221.0$ kJ
$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ $\quad \Delta_f H° = -483.6$ kJ

3. Calculate the standard enthalpy change for reaction (2)—the combustion of coal gas—from the following information.

$2C(s) + O_2(g) \rightarrow 2CO(g)$ $\quad \Delta_f H° = -221.0$ kJ
$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta_f H° = -393.5$ kJ
$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ $\quad \Delta_f H° = -483.6$ kJ

4. Calculate the standard enthalpy change for reaction (3)—the methanation of coal gas—using the following information.

$2H_2(g) + O_2(g) \rightarrow 2H_2O(g)$ $\quad \Delta_f H° = -483.6$ kJ
$2C(s) + O_2(g) \rightarrow 2CO(g)$ $\quad \Delta_f H° = -221.0$ kJ
$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2H_2O(g)$ $\quad \Delta_c H° = -802.7$ kJ
$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta_f H° = -393.5$ kJ

DID YOU KNOW ?

Coal Gasification

The gasification of coal has three potential applications. Besides the electric power generation described below, there is the possibility of using steam to gasify deep coal to bring the energy to the surface. Another possibility is to use the excess coke produced by oil sands refining to produce electric power or methane.

DID YOU KNOW ?

Cogeneration

Cogeneration, also called CHP (combined heat and power), is a modern technology for the simultaneous production of electricity and heat using a single fuel. The heat is sometimes used to generate more electricity. Cogeneration provides greater efficiencies than traditional generating methods. The process illustrated in Figure 5 is one example of cogeneration.

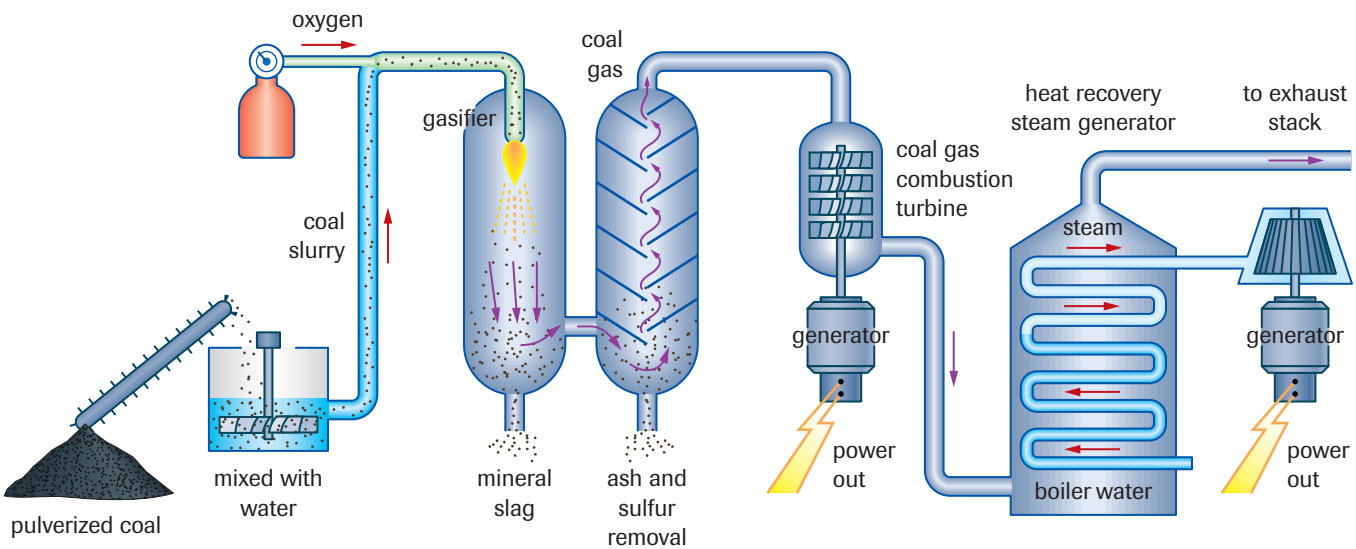


Figure 5
Electric power-generating stations that use coal as a fuel are 30% to 40% efficient. Coal gasification and combustion of the coal gas provide one alternative to burning coal. Efficiency is improved by about 10% by using both a combustion turbine and a steam turbine to produce electricity.

WEB Activity

Simulation—Hess' Law

Titantium(IV) chloride is an unusual compound: it is a liquid metal halide. Chemists explain this anomaly by theorizing that it is a molecular substance, rather than ionic. It is also very corrosive, forming hydrochloric acid when it comes into contact with water, so is only used with very strict safety precautions. We can, however, discover some of its properties by using simulations.

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LAB EXERCISE 11.B

Testing Hess' Law

The following data are from a test of Hess' law using a calorimeter. Use these data in your prediction, assuming that combustion produces carbon dioxide gas and liquid water.

$5\,C(s) + 6\,H_2(g) \rightarrow C_5H_{12}(l)$ $\quad \Delta_f H° = -173.5$ kJ

$C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta_f H° = -393.5$ kJ

$H_2(g) + \frac{1}{2}\,O_2(g) \rightarrow H_2O(g)$ $\quad \Delta_f H° = -241.8$ kJ

$H_2O(l) \rightarrow H_2O(g)$ $\quad \Delta_{vap} H° = +40.65$ kJ

Then complete the Prediction, Analysis, and Evaluation of the investigation report.

Purpose

The purpose of this investigation is to test Hess' law.

Problem

What is the standard molar enthalpy of combustion of pentane?

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (2, 3)
○ Hypothesis	○ Procedure	
● Prediction	○ Evidence	

Design

Hess' law is used to predict the standard molar enthalpy of combustion of pentane. To test the prediction and the acceptability of the law, the standard molar enthalpy of combustion of pentane is determined calorimetrically.

Evidence

mass of pentane reacted = 2.15 g
mass of water equivalent to calorimeter = 1.24 kg
initial temperature of calorimeter and contents = 18.4 °C
final temperature of calorimeter and contents = 37.6 °C

(Note that the calorimeter has the same thermal energy gain or loss as the water equivalent stated.)

LAB EXERCISE 11.C

Analysis Using Hess' Law

Most natural gas is burned as fuel to provide heat. However, natural gas is also a source of hydrogen gas for producing ammonia-based fertilizers. Complete one possible Analysis for the investigation report. Not all the evidence need be used for a particular analysis.

Purpose

The purpose of this exercise is to use Hess' law to determine an enthalpy change.

Problem

What is the standard enthalpy change for the production of hydrogen from methane and steam?

$CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\,H_2(g)$ $\quad \Delta_r H° = ?$

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	○ Evaluation
○ Hypothesis	○ Procedure	
○ Prediction	○ Evidence	

Evidence

(1) $2\,C(s) + O_2(g) \rightarrow 2\,CO(g)$ $\quad \Delta_c H° = -221.0$ kJ

(2) $CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$ $\quad \Delta_c H° = -802.5$ kJ

(3) $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$ $\quad \Delta_r H° = -41.2$ kJ

(4) $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g)$ $\quad \Delta_c H° = -483.6$ kJ

(5) $C(s) + 2\,H_2(g) \rightarrow CH_4(g)$ $\quad \Delta_f H° = -74.6$ kJ

(6) $C(s) + H_2O(g) \rightarrow CO(g) + H_2(g)$ $\quad \Delta_r H° = +131.3$ kJ

(7) $2\,CO(g) + O_2(g) \rightarrow 2\,CO_2(g)$ $\quad \Delta_c H° = -566.0$ kJ

(8) $CO(g) + H_2(g) + O_2(g) \rightarrow CO_2(g) + H_2O(g)$ $\quad \Delta_r H° = -524.8$ kJ

INVESTIGATION 11.3 *Introduction*

Applying Hess' Law

Magnesium burns rapidly, releasing heat and light. The enthalpy change of this reaction can be calculated from measurements involving a highly sophisticated calorimeter, but not a polystyrene cup calorimeter. The enthalpy change for the combustion of magnesium can, however, be determined by applying Hess' law to a series of chemical reaction equations (see page 517).

Purpose

The purpose of this investigation is to use Hess' law to determine a molar enthalpy of combustion.

Problem

What is the standard molar enthalpy of combustion for magnesium?

Prediction

The standard molar enthalpy of combustion for magnesium is −601.6 kJ/mol, as per the table of standard molar enthalpies of formation (Appendix I). (The standard molar enthalpy of combustion of magnesium is the same as the standard molar enthalpy of formation of magnesium oxide because both processes have the same chemical reaction equation.)

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
○ Prediction	● Evidence	

Design

The enthalpy changes for the first two reactions in the series (on page 517) are determined empirically using a polystyrene calorimeter, and the third enthalpy change is obtained from Appendix I. The three $\Delta_f H$ values are used, along with Hess' law, to obtain the molar enthalpy of combustion for magnesium.

To perform this investigation, turn to page 517.

EXPLORE *an issue*

Alternative Energy Sources and Technologies

Issue Checklist

○ Issue	○ Design	● Analysis
○ Resolution	● Evidence	● Evaluation

Photosynthesis is the key to life on Earth. For millions of years green plants have used sunlight, water, and carbon dioxide to produce biomass. When prehistoric plants and the tiny animals that ate them died, the organic materials they contained settled on the bottom of swamps, lakes, and seas, where rocky sediments eventually buried them. Fossil fuels (coal, oil, and natural gas) are the remains of these ancient organisms, transformed by a combination of heat and pressure. While living organisms are composed mainly of carbon, hydrogen, and oxygen, the process of being converted into fossil fuels removed the oxygen from the organic matter so the remaining compounds are referred to as *hydrocarbons*. Although Alberta's hydrocarbon reserves are very large, they are also finite because of the time and conditions necessary to produce fossil fuels (**Figure 6**).

For several decades researchers have been looking for alternatives to fossil fuels. One promising alternative involves using "renewable" plant material from managed forests or farms, which takes a few months or years to grow, rather than using fossil fuels, which take about 300 million years to form. Another option takes waste agricultural or industrial material.

A bio-based economy uses renewable energy and renewable materials instead of fossil fuels to provide the energy, chemicals, and material for society. Proponents of a bio-based economy see a fuel source in the waste carbon compounds produced by agriculture and forestry (**Figure 7**). These waste plant and animal materials could supply the biomass to produce ethanol or biodiesel, or even hydrogen gas

Figure 6
A Suncor Energy Inc. shovel fills the bed of a heavy hauler truck in the oil sands region near Fort McMurray.

Figure 7
The new Grande Prairie EcoPower™ Centre uses wood waste to co-generate both electricity and steam for use in two nearby sawmills. The Government of Alberta and the City of Grande Prairie have committed to a long-term purchase of electricity from the EcoPower™ Centre.

for use in fuel cells. This is not a new concept. Wood has been and is being used for heating and cooking. When wood biomass was unavailable, people made use of other biomass sources. For example, prairie First Nations people used buffalo dung for heating, and the Inuit have made efficient use of seal oil for heating and light.

Biomass could also produce the feedstock chemicals used to make the wide array of synthetic materials, including plastics, that are currently synthesized from petrochemicals. Proponents insist that such an economy has the potential to be both renewable and sustainable.

Many of the processes needed to utilize the biomass from agriculture and forestry are already in place, but would need to be adapted and expanded. For example, ethanol is now made largely from starch, which could be used as a food source for people or animals. It may be more efficient to convert the waste parts of crop plants, or the waste from sawmills, into ethanol. Similarly, the technology exists to turn waste oil from the fast-food industry into diesel fuel. The source of biomass could be increased by planting crops specifically for that purpose, using farmland that is not currently being utilized. Some researchers estimate that we could obtain up to 25% of our current energy demands by converting waste biomass into useable fuels.

Many people in the energy industry are skeptical about the practicality of a bio-based economy. They argue that the technology for finding and using fossil fuels is already well established, while the technologies of a bio-based economy will take considerable time and money to develop. In their view, research would be better directed at finding cleaner and more efficient ways of extracting and using traditional fossil fuels. For example, three projects aimed at developing cleaner fossil fuels are

- enhanced coal bed methane
- heavy oil recovery using vapour extraction
- paste technology for oil sands tailings treatment

Energy issues have pros and cons from many perspectives (see Appendix D) .Your minimum task is to write an energy Issue with a technology emphasis and complete the Resolution and Design. (An example is provided.) Your teacher may also give you instructions to complete the Evidence, Analysis, and Evaluation components of this issue resolution process.

Issue

What is the most effective use of government funding for research in energy technology?

Resolution

The provincial and federal governments should direct all their research funding to bio-based energy technology.

Design

Within small groups, research the pros and cons of using public money to fund bio-based versus fossil-fuel-based energy technology.

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Section 11.4 Questions

Use the following information to answer questions 1 to 5.

Energy may be produced or consumed by the production and synthesis of fuels such as alcohols. Use Hess' law to determine the quantity of energy required or released during the production of various alcohol fuels.

1. Ethanol (grain alcohol) is a renewable energy gasoline additive that can be made from glucose obtained from the fermentation of grain. Determine the standard enthalpy of reaction for

$C_6H_{12}O_6(s) \rightarrow 2\,C_2H_5OH(l) + 2\,CO_2(g) \qquad \Delta_r H = ?$

from the following reaction equations.

$C_6H_{12}O_6(s) + 6\,O_2(g) \rightarrow 6\,CO_2(g) + 6\,H_2O(l) + 2\,803.1\ kJ$

$C_2H_5OH\ (l) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 3\,H_2O(l) + 1\,366.8\ kJ$

2. Ethanol can also be made from the nonrenewable resource ethene. First, ethane from natural gas is cracked to produce ethene, as is done in several ethene plants in Alberta.

$C_2H_6(g) \rightarrow C_2H_4(g) + H_2(g) \qquad \Delta_r H = ?$

The second step is the production (synthesis) of ethanol from ethene by an addition reaction with water.

$C_2H_4(g) + H_2O(l) \rightarrow C_2H_5OH(l) \qquad \Delta_r H = ?$

(a) Using the following enthalpies of reaction, predict the enthalpy of reaction for the ethane cracking reaction equation given.

$2\,C_2H_6(g) + 7\,O_2(g) \rightarrow 4\,CO_2(g) + 6\,H_2O(l) + 3\,120.8\ kJ$

$C_2H_4(g) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(l) \qquad \Delta_c H = -1\,411.0\ kJ$

$H_2(g) + \frac{1}{2}\,O_2(g) \rightarrow H_2O(l) \qquad \Delta_c H = -285.8\ kJ$

(b) Using the following enthalpies of reaction, predict the enthalpy of reaction for the addition reaction of ethene to water to produce ethanol.

$C_2H_4(g) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 2\,H_2O(l) \qquad \Delta_c H^\circ = -1\,411.0\ kJ$

$C_2H_5OH\ (l) + 3\,O_2(g) \rightarrow 2\,CO_2(g) + 3\,H_2O(l) + 1\,366.8\ kJ$

(c) If the reaction equations that you just used (to determine the enthalpies of reaction for the cracking of ethane and for the addition of water to ethene) are not the actual steps used in these chemical processes, then why does this Hess' law technique work?

3. Methanol (wood alcohol) is a fuel that can be made from biomass (such as cellulose obtained from aspen trees). Methanol can also be made from natural gas. The overall synthesis reaction for the latter case is

$CH_4(g) + H_2O(l) \rightarrow CH_3OH(l) + H_2(g)$ $\quad \Delta_r H° = ?$

Use the reactions below and Hess' law to predict the standard molar enthalpy of synthesis of methanol by this process.

$CO(g) + 3\,H_2(g) \rightarrow CH_4(g) + H_2O(l)$ $\quad \Delta_r H° = -249.9$ kJ

$CO(g) + 2\,H_2(g) \rightarrow CH_3OH(l)$ $\quad \Delta_r H° = -128.7$ kJ

4. Consider each of the net reactions in questions 1, 2, and 3 for the synthesis of ethanol and methanol fuels: ethanol from glucose, ethanol from ethane, and methanol from methane. Create a summary, such as a table, in which to present your answers to the following questions.

(a) Is each net synthesis reaction endothermic or exothermic?

(b) Reference (look up) and record the molar enthalpies of combustion of each fuel.

(c) Compare the molar enthalpy of synthesis of each fuel (ethanol and methanol) to the molar enthalpy of combustion of the fuel. Is the net chemical process, considering the synthesis and the combustion of the fuel, endothermic or exothermic?

(d) Quantitatively, what is the net standard molar enthalpy of reaction for the combined chemical process of synthesis and combustion for each fuel in each of the synthesis processes?

(e) What implications does this have for the production of these alcohols as fuels?

5. Of course, the net energy is less than that calculated in the previous question. What other kinds of energy costs might be associated with the manufacture of fuels such as ethanol and methanol?

---

6. Describe some other technological solutions to the problem of our growing per-capita and per-country demand for energy.

7. Ethyne gas may react with hydrogen gas to form ethane gas in the following reaction:

$C_2H_2(g) + 2\,H_2(g) \rightarrow C_2H_6(g)$

Predict the standard enthalpy change for the reaction of 200 g of ethyne, using the following information. After using Hess' law, refer back to Sample Problem 11.2 on page 491.

(1) $C_2H_2(g) + \frac{5}{2}\,O_2(g) \rightarrow 2\,CO_2(g) + H_2O(l)$ $\quad \Delta_1 H° = -1299$ kJ

(2) $H_2(g) + \frac{1}{2}\,O_2(g) \rightarrow H_2O(l)$ $\quad \Delta_2 H° = -286$ kJ

(3) $C_2H_6(g) + \frac{7}{2}\,O_2(g) \rightarrow 2\,CO_2(g) + 3\,H_2O(l)$ $\quad \Delta_3 H° = -1560$ kJ

8. As an alternative to combustion, coal gas can undergo a process called *methanation*.

$3\,H_2(g) + CO(g) \rightarrow CH_4(g) + H_2O(g)$ $\quad \Delta_r H° = ?$

Determine the standard enthalpy change involved in the reaction of 300 g of carbon monoxide in this methanation reaction.

9. $C(s) + O_2(g) \rightarrow CO_2(g)$ $\quad \Delta_c H° = -393.5$ kJ

$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g) + 483.6$ kJ

$\Delta_c H°_m = -5074.1$ kJ/mol

C_8H_{18}

Use Hess' law and combustion reaction information above to predict the standard enthalpy change for the formation of octane from its elements.

$8\,C(s) + 9\,H_2(g) \rightarrow C_8H_{18}(l)$ $\quad \Delta_f H° = ?$

10. Bacteria sour wines and beers by converting ethanol, C_2H_5OH, into acetic acid, CH_3COOH. Assume that the reaction is

$C_2H_5OH(l) + O_2(g) \rightarrow CH_3COOH(l) + H_2O(l)$

The standard molar enthalpies of combustion of ethanol and acetic acid to produce liquid water are, respectively, −1367 kJ/mol and −875 kJ/mol. Write chemical equations for the combustions, and use Hess' law to determine the standard enthalpy change for the conversion of ethanol to acetic acid.

11. A series of calorimetric experiments is performed to test Hess' law. Complete the Purpose, Prediction, Analysis, and Evaluation (2, 3) sections of the investigation report. Evaluation should include a calculation of a percent difference in the experiment, given that the $\Delta_r H$ for equation (1) should equal the sum of of the $\Delta_r H$'s for equations (2) and (3).

Problem

What is the enthalpy change for the reaction of aqueous potassium hydroxide with hydrobromic acid?

Design

Three calorimetry experiments are performed: Experiment 1 for the Prediction and Experiments 2 and 3 for the Analysis. The three thermochemical reactions are

(1) $HBr(aq) + KOH(aq) \rightarrow H_2O(l) + KBr(aq)$ $\quad \Delta_1 H = ?$ kJ

(2) $KOH(s) \rightarrow KOH(aq)$ $\quad \Delta_2 H = ?$ kJ

(3) $KOH(s) + HBr(aq) \rightarrow H_2O(l) + KBr(aq)$ $\quad \Delta_3 H = ?$ kJ

Evidence

Table 1 Reaction of KOH(aq) and HBr(aq)

Observation	Experiment 1	Experiment 2	Experiment 3
quantity of reactant 1	100.0 mL of 1.00 mol/L HBr(aq)	5.61 g KOH(s) in 200.0 mL of solution	5.61 g KOH(s)
quantity of reactant 2	100.0 mL of 1.00 mol/L KOH(aq)	N/A	200.0 mL of 0.50 mol/L HBr(aq)
initial temperature	20.0 °C	20.0 °C	20.0 °C
final temperature	22.5 °C	24.1 °C	26.7 °C

11.5 Molar Enthalpies of Formation

Figure 1
As a cyclist coasts downhill, his potential energy decreases—it is converted to kinetic energy.

Chemists rely on conventions to simplify explanations and communication. For example, SATP is a set of internationally accepted conditions that define a *standard state* for formation reactions. It is convenient to set at zero the value for the enthalpies of elements in their most stable form at SATP (the standard conditions for elements). This convention, defining elements as the reference point at which the potential energy is zero, is the **reference energy state**. This does not mean that the enthalpy of an element is *always* considered to be zero; in another situation, a different convention might be more convenient. (Similarly, the Celsius temperature scale sets 0 °C at the freezing point of pure water. This is a convenient reference point but it does not mean that water molecules have zero thermal or kinetic energy at that temperature.)

A formation reaction is a reaction in which a compound is formed from its constituent elements. The enthalpy change calculated from measurements of a formation reaction under standard conditions is called the **standard enthalpy of formation**, $\Delta_f H°$. We can now describe $\Delta_f H°$ theoretically: as a change in enthalpy from zero (the reference enthalpy of formation of the elements) to some final value determined by the enthalpy change. For example,

$$H_2(g) + \tfrac{1}{2}\,O_2(g) \rightarrow H_2O(l) \qquad \Delta_f H° = -285.8 \text{ kJ}$$

$$\Delta_f H° \text{ (kJ)}: \quad 0 \qquad\quad 0 \qquad\quad -285.8$$

The potential energy decreases from 0 kJ for the reactants to –285.8 kJ for the product. In other words, the reactants are at a higher chemical potential energy than the product. This potential energy is transferred to the surroundings and appears as thermal energy or other forms of energy. Suppose you were seated on a bicycle at the top of a hill and coasted downhill. Your potential energy at the top of the hill is converted into kinetic energy as you move from a point of higher potential energy (top) to one of lower potential energy (bottom) (**Figure 1**). Of course, to return to the top of the hill you must supply the energy to move to a higher potential energy. Similarly, to convert the water back into hydrogen and oxygen requires that energy be added, specifically 285.8 kJ/mol of water.

We can use tables of standard molar enthalpies of formation (Appendix I) to compare the stabilities of compounds. Most compounds are formed exothermically from their elements, and the molar enthalpies of formation are negative. This means that the compounds are more stable than their elements (defined at 0 kJ/mol). **Thermal stability** is the tendency of a compound to resist decomposition when heated. The standard molar enthalpies of formation give an indication of thermal stability. The lower the value, the greater the thermal stability. By analogy, you are physically more stable when you are at a lower gravitational potential energy—closer to the ground. Another example is a comparison of the standard molar enthalpies of formation for tin(II) oxide and tin(IV) oxide, which can be obtained in Appendix I.

$$\Delta_f H_m° = -280.7 \text{ kJ/mol} \quad (SnO) \qquad\qquad \Delta_f H_m° = -577.6 \text{ kJ/mol} \quad (SnO_2)$$

Chemists explain that tin(IV) oxide has more thermal stability than tin(II) oxide because tin(IV) oxide has a more negative molar enthalpy of formation. The lower the energy, the greater the stability.

Enthalpies of reaction and molar enthalpies of reaction for many chemical reactions can be predicted from the molar enthalpies of formation of elements and compounds. Sample Problem 11.5 illustrates the derivation of the molar enthalpies of formation method from the Hess' law method.

DID YOU KNOW ?

Zero Energy

Setting a reference point for the energy (especially potential energy) being zero is common practice. Examples of "zero points" include

- zero (0 °C) on the Celsius scale
- zero gravitational potential energy at the surface of Earth
- zero electrical potential energy for the ground
- zero electrode potential for the hydrogen electrode (see Chapter 13)
- zero enthalpy of formation for elements under standard conditions

Ideally, these reference points are set (and sometimes changed) by international agreement (for example, by IUPAC) to facilitate international communication.

SAMPLE problem 11.5

How can molar enthalpies of formation be used to calculate enthalpies of reaction? Consider the slaking of lime, calcium oxide, represented by the following chemical reaction equation (**Figure 2**).

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s) \qquad \Delta_r H° = ?$$

What is the standard enthalpy of reaction for this reaction?

First, write the formation equation and corresponding standard enthalpy of formation (Appendix I) for each compound in the given equation. To find the enthalpy of formation, $\Delta_f H°$, multiply the chemical amount by the molar enthalpy of formation, $\Delta_f H° = n\,\Delta_f H_m°$.

$$Ca(s) + \tfrac{1}{2}\,O_2(g) \rightarrow CaO(s) \qquad \Delta_f H° = 1\text{ mol} \times -634.9\text{ kJ/mol}$$

$$H_2(g) + \tfrac{1}{2}\,O_2(g) \rightarrow H_2O(l) \qquad \Delta_f H° = 1\text{ mol} \times -285.8\text{ kJ/mol}$$

$$Ca(s) + O_2(g) + H_2(g) \rightarrow Ca(OH)_2(s) \qquad \Delta_f H° = 1\text{ mol} \times -986.1\text{ kJ/mol}$$

By adding the third equation to the reverse of the first two equations, the chemical equation required for the slaking of lime is obtained.

$$CaO(s) \rightarrow \cancel{Ca(s)} + \cancel{\tfrac{1}{2}\,O_2(g)} \qquad -\Delta_f H°$$

$$H_2O(l) \rightarrow \cancel{H_2(g)} + \cancel{\tfrac{1}{2}\,O_2(g)} \qquad -\Delta_f H°$$

$$\cancel{Ca(s)} + \cancel{O_2(g)} + \cancel{H_2(g)} \rightarrow Ca(OH)_2(s) \qquad \Delta_f H°$$

$$\text{Net: } CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s) \qquad \Delta_r H° = ?$$

Applying Hess' law gives the following equation:

$$\Delta_r H° = \underset{Ca(OH)_2}{\Delta_f H°} + (-\underset{CaO}{\Delta_f H°}) + (-\underset{H_2O}{\Delta_f H°})$$

Notice that the net enthalpy change is equal to the enthalpy of formation for the product minus the enthalpies of formation of the reactants.

$$\Delta_r H° = \underset{Ca(OH)_2}{\Delta_f H°} - (\underset{CaO}{\Delta_f H°} + \underset{H_2O}{\Delta_f H°})$$

Substituting the definition $\Delta_f H° = n\,\Delta_f H_m°$ (Section 11.1) and combining terms results in the following formula, where $\Sigma n\,\Delta_{fP} H_m°$ is the standard enthalpy change of all the products (P), and $\Sigma n\,\Delta_{fR} H_m°$ is the standard enthalpy change of all the reactants (R).

$$\Delta_r H° = \underset{Ca(OH)_2}{\Sigma n\,\Delta_{fP} H_m°} - \underset{CaO + H_2O}{\Sigma n\,\Delta_{fR} H_m°}$$

$$= \underset{Ca(OH)_2}{n\,\Delta_{fP} H_m°} - (\underset{CaO}{n\,\Delta_{fR} H_m°} + \underset{H_2O}{n\,\Delta_{fR} H_m°})$$

$$= \left(1\text{ mol } Ca(OH)_2 \times \frac{-986.1\text{ kJ}}{1\text{ mol } Ca(OH)_2}\right) - \left(1\text{ mol CaO} \times \frac{-634.9\text{ kJ}}{1\text{ mol CaO}} + 1\text{ mol } H_2O \times \frac{-285.8\text{ kJ}}{1\text{ mol } H_2O}\right)$$

$$= -986.1\text{ kJ} - (-634.9 + (-285.8))\text{ kJ}$$

$$= -65.4\text{ kJ}$$

According to Hess' law and empirically determined standard molar enthalpies of formation, the standard enthalpy change for the slaking of lime is reported as follows:

$$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s) \qquad \Delta_r H° = -65.4\text{ kJ}$$

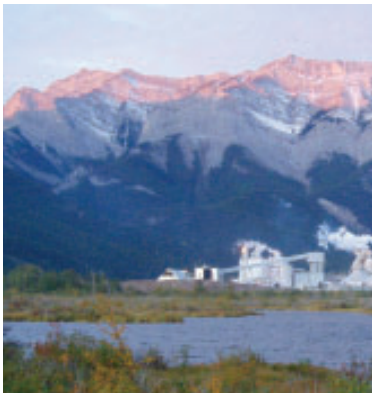
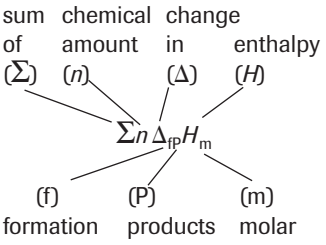


Figure 2
Limestone, in this case from the mountains near Exshaw, Alberta, can be decomposed to lime. The lime is then converted to slaked lime. The limestone, lime, and slaked lime all have technological applications.

Learning Tip

The symbols for this method of calculating the enthalpy of reaction are meant to help but, initially, can be confusing.



This set of symbols translates into: "the sum of the chemical amounts times the molar enthalpies of formation of the products." This set of symbols changes from P for the products to R for the reactants. Learn the concept (that is, the change in enthalpy is the energy of the products minus the energy of the reactants), and the equation then makes sense. See the Summary on the next page.

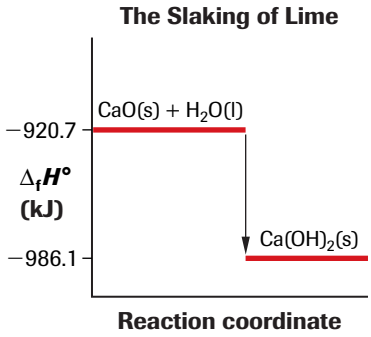


Figure 3
Enthalpy change diagram for the slaking of lime. The two summation (Σ) terms in the mathematical formula become the positions of the reactants and the products on the potential energy scale.

The enthalpy change of the reaction in Sample Problem 11.5 can be theoretically described by an enthalpy change diagram similar to a chemical potential energy diagram (**Figure 3**). Note that the derived formula provides a negative enthalpy change for an exothermic reaction, consistent with the accepted convention. Also note that the value of $\Sigma n_{fP} H_m°$ appears on the diagram as the standard enthalpy of formation of the product, and the value of $\Sigma n_{fR} H_m°$ is the standard enthalpy of formation of the reactants. The label on the *y*-axis must be "standard enthalpy of formation, $\Delta_f H°$."

SUMMARY Using Enthalpies of Formation to Predict Standard Enthalpy Changes, $\Delta_r H°$

According to Hess' law, the net enthalpy change for a chemical reaction is equal to the sum of the chemical amounts times the molar enthalpies of formation of the products minus the sum of the chemical amounts times the molar enthalpies of formation of the reactants (**Figure 4**).

$$\Delta_r H° = \Sigma n \Delta_{fP} H_m° - \Sigma n \Delta_{fR} H_m°$$

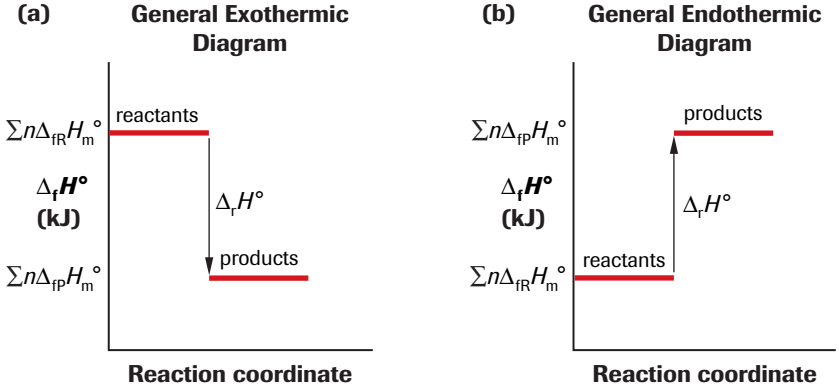


Figure 4
The two summation terms in the net enthalpy change formula provide the location of the reactants and products on enthalpy change diagrams.

Learning Tip

Think of the fairly complicated equation for calculating the enthalpy of reaction

$\Delta_c H° = \Sigma n \Delta_{fP} H_m° - \Sigma n \Delta_{fR} H_m°$

as communicating a necessary concept. Basically, the equation communicates the idea that the change in enthalpy for a reaction is the total chemical potential energy of the products minus that of the reactants. The order of "products minus reactants" makes the sign for the enthalpy change come out correctly.

COMMUNICATION example

Methane is burned in furnaces and in some power plants (**Figure 5**). What is the standard molar enthalpy of combustion of methane? Assume that water vapour is a product.

Solution

$$CH_4(g) + 2\,O_2(g) \rightarrow CO_2(g) + 2\,H_2O(g)$$

$$\Delta_c H° = \left(1\text{ mol } CO_2 \times \frac{-393.5\text{ kJ}}{1\text{ mol } CO_2} + 2\text{ mol } H_2O \times \frac{-241.8\text{ kJ}}{1\text{ mol } H_2O}\right) - \left(1\text{ mol } CH_4 \times \frac{-74.6\text{ kJ}}{1\text{ mol } CH_4} + 2\text{ mol } O_2 \times \frac{0\text{ kJ}}{1\text{ mol } O_2}\right)$$

$$= -877.1\text{ kJ} - (-74.6\text{ kJ})$$

$$= -802.5\text{ kJ}$$

$$\Delta_c H°_{CH_4} = \frac{-802.5\text{ kJ}}{1\text{ mol}} = -802.5\text{ kJ/mol}$$

According to the molar enthalpy of formation method, the standard molar enthalpy of combustion of methane to produce water vapour is −802.5 kJ/mol.

The combustion of methane is communicated as an enthalpy change diagram in **Figure 6.** Note that the labelling of the *y*-axis is different from that in a chemical potential energy diagram.

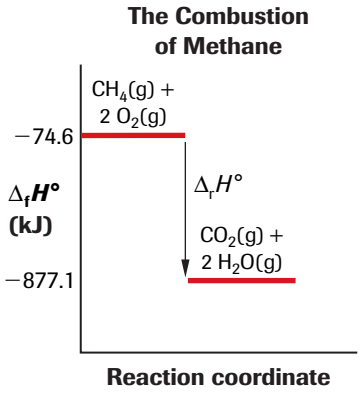


Figure 6
The enthalpy change diagram for the combustion of methane conveys the same information as the equation in the Communication Example.

Figure 5
A blue flame in a home furnace indicates that the gas is burning completely, and no carbon monoxide gas is being produced.

LAB EXERCISE 11.D

Testing $\Delta_r H°$ from Formation Data

The molar enthalpies of formation method was created earlier in this section from Hess' law. Both problem-solving methods are consistent with the law of conservation of energy. This lab exercise tests this new method against evidence gathered from calorimetry.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (2, 3)
○ Hypothesis	○ Procedure	
● Prediction	○ Evidence	

Purpose

The purpose of this problem is to test the use of molar enthalpies of formation as a method of predicting the enthalpy change of a reaction.

Problem

What is the standard molar enthalpy of combustion of methanol?

Design

Methanol is burned in excess oxygen in a special calorimeter. Assume that liquid water is produced.

Evidence

mass of methanol reacted = 4.38 g
mass of water equivalent to calorimeter = 2.60 kg
initial temperature of calorimeter = 20.4 °C
final temperature of calorimeter = 27.9 °C

(Note that the calorimeter has the same thermal energy gain or loss as the water equivalent stated.)

WEB Activity

Web Quest—Rocket Fuel Thermochemistry

This Web Quest allows you to work as part of a research team, exploring the differences between types of rocket fuels. You can use your results to support a competitor for the latest $10 million Ansari X Prize, given for inventions that promote space travel. You can look at the properties of reactants and products as well as the energy changes involved in the rocket fuel reactions. You and your team can create and deliver a PowerPoint™ presentation about your findings.

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Section 11.5 Questions

1. Which of the following fuels has the greater thermal stability: methanol or ethanol? Provide your reasoning.
2. Methane, the major component of natural gas, is used as a source of hydrogen gas to produce ammonia. Ammonia is used as a fertilizer and a refrigerant, and is used to manufacture fertilizers, plastics, cleaning agents, and prescription drugs. The following questions refer to some of the chemicals involved.
   (a) The first step in the production of ammonia is the reaction of methane with steam using a nickel catalyst. Use the molar enthalpies of formation method to predict the $\Delta_r H°$ for the following reaction:
   $CH_4(g) + H_2O(g) \rightarrow CO(g) + 3\,H_2(g)$
   (b) Another step of this process is the further reaction of carbon monoxide to produce more hydrogen. Iron and zinc–copper catalysts are used. Predict the $\Delta_r H°$.
   $CO(g) + H_2O(g) \rightarrow CO_2(g) + H_2(g)$
   (c) After the carbon dioxide gas is removed, the hydrogen reacts with nitrogen obtained from the air. Predict the $\Delta_f H°$ to form 2 mol of ammonia.
   $N_2(g) + 3\,H_2(g) \rightarrow 2\,NH_3(g) \qquad \Delta_f H° = ?$
3. Nitric acid, required in the production of nitrate fertilizers, is produced from ammonia by the Ostwald process (**Figure 7**). Predict the standard enthalpy change for each reaction in the process, as written, and then predict the standard molar enthalpy of reaction for the first reactant listed in each equation.
   (a) $4\,NH_3(g) + 5\,O_2(g) \rightarrow 4\,NO(g) + 6\,H_2O(g)$
   (b) $2\,NO(g) + O_2(g) \rightarrow 2\,NO_2(g)$
   (c) $3\,NO_2(g) + H_2O(l) \rightarrow 2\,HNO_3(l) + NO(g)$

Figure 7
An Ostwald process plant converts ammonia to nitric acid cleanly and efficiently. Unreacted gases and energy from the exothermic reactions are recycled. Catalytic combustors burn noxious fumes to minimize environmental effects and to supply additional energy to operate the plant.

4. Ammonium nitrate fertilizer is produced by the reaction of ammonia with the nitric acid resulting from the series of reactions given in question 3. Ammonium nitrate is one of the most important fertilizers for increasing crop yields.
   (a) Predict the standard enthalpy change of the reaction used to produce ammonium nitrate.
   $NH_3(g) + HNO_3(l) \rightarrow NH_4NO_3(s) \qquad \Delta_r H° = ?$
   (b) Sketch an enthalpy change diagram for the reaction of ammonia and nitric acid.
5. During World War II an oil embargo on Germany left that country short of fuels. Germany responded by producing and burning ammonia as a fuel for cars, trucks, and tanks.
   (a) What is the standard molar enthalpy of combustion of ammonia gas? Assume that nitrogen dioxide and liquid water are the products of combustion.
   (b) Draw an enthalpy change diagram for the burning of 4 mol of ammonia, including labelling the *y*-axis with the energy involved.
   (c) Evaluate this technological solution from two perspectives.
6. The refining of sour natural gas removes toxic hydrogen sulfide from the natural gas at many of the gas plants in Alberta. One of the reactions during the conversion of the hydrogen sulfide to sulfur is the combustion of some of this gas.
   (a) What is the standard molar enthalpy of combustion of hydrogen sulfide to produce sulfur dioxide gas and liquid water?
   (b) Communicate the standard enthalpy of combustion of hydrogen sulfide in the four ways described earlier in this chapter.
   (c) Should the products of the combustion of hydrogen sulfide be released into the atmosphere? Provide your reasoning.
7. You now have several ways of knowing the enthalpy of a chemical reaction. The simplest are to be given, to reference, or to memorize the enthalpy change for selected chemical reactions. The law of conservation of energy has so far provided you with three other methods for calculating an enthalpy change. How would you recognize when you are supposed to use each of these methods, listed below, in a question? That is, what would the question look like? What information would you be given in the question?
   (a) calorimetry
   (b) Hess' law
   (c) molar enthalpies of formation

Use this information to answer questions 8 and 9.

There is a whitish-grey mountain near Exshaw, Alberta from which limestone is mined (Figure 2, page 511). This limestone is then thermally decomposed into lime and carbon dioxide. Some of the lime is then slaked: the lime reacts with water to produce slaked lime. The lime is used in a complex cement mixture to create concrete. The lime can also be used to make mortar for bricklaying.

8. (a) Use standard molar enthalpies of formation to predict the standard reaction enthalpy for the first step of the chemical process to produce slaked lime.

$CaCO_3(s) \rightarrow CaO(s) + CO_2(g) \qquad \Delta_r H° = ?$

(b) Now predict the $\Delta_r H°$ for the second step.

$CaO(s) + H_2O(l) \rightarrow Ca(OH)_2(s) \qquad \Delta_r H° = ?$

(c) Write the overall (net) reaction equation for the two steps, and use Hess' law to predict the net standard enthalpy of reaction.

(d) Use the net reaction equation and molar enthalpies of formation to predict the net standard reaction enthalpy.

(e) Compare your answers in (c) and (d).

(f) List some further technological applications of lime and slaked lime.

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9. Slaked lime is used in mortar, for building brick walls (**Figure 8**). As the wall "cures," the mortar is strengthened by the reaction of slaked lime with carbon dioxide from the air to produce calcium carbonate and water. Look back at your answers for the previous question to determine the standard enthalpy of reaction without having to use Hess' law or standard molar enthalpies of reaction.

Figure 8
The rate at which mortar cures can affect its strength and longevity.

Extension

10. The prediction of enthalpy changes using standard molar enthalpies of formation depends entirely on the availability of tables of standard molar enthalpies of formation. Many of these $\Delta_f H_m°$ values can be initially determined from $\Delta_c H_m°$ values by using the formation method equation. In this problem you will use a known $\Delta_c H_m°$ value to calculate a corresponding $\Delta_f H_m°$ value. Complete the Analysis of the investigation report. Work out your own problem-solving approach here.

Problem
What is the standard molar enthalpy of formation for hexane, $C_6H_{14}(l)$?

Design
The standard molar enthalpy of formation for hexane is determined using the concept
$\Delta_c H° = \Sigma n \Delta_{fp} H_m° - \Sigma n \Delta_{fr} H_m°$
to describe the hexane combustion reaction. The standard molar enthalpy of combustion of hexane is obtained from a reference source.

Evidence
The CRC Handbook of Chemistry and Physics lists the calorimetric value for the standard molar enthalpy of combustion for hexane as −4 162.9 kJ/mol. Carbon dioxide gas and liquid water are the only products of the combustion.

11. Besides specific heat capacity (e.g., 4.19 J/(g•°C) for water), there are other energy descriptions for chemical substances. Based upon the units used for these quantities, describe as best you can what the quantity is about and where it might be of importance or of use.

(a) volumetric heat capacity (e.g., 1.2 kJ/(m^3•°C) for air at SATP)

(b) molar heat capacity (e.g., 75.5 kJ/(mol•°C) for water)

(c) heat capacity (e.g., 9.12 kJ/°C for a particular bomb calorimeter)

(d) specific heat (e.g., 13.4 kJ/g of water for the formation of water vapour from hydrogen and oxygen)

12. The bomb calorimeter and flow calorimeter are high-tech calorimeters used to determine, for example, the energy value of foodstuffs. In a test of the molar enthalpies of formation method, a Prediction is made of the standard molar enthalpy of respiration of glucose. Use calorimetry in the Analysis to obtain an empirical test value. Complete the Evaluation (2 and 3). (Assume liquid water is produced.)

Problem
What is the standard molar enthalpy of respiration of glucose?

Evidence
mass of glucose = 3.00 g
heat capacity of calorimeter = 8.52 kJ/°C
initial temperature = 25.02 °C
final temperature = 30.49 °C

13. There are many commercial products sold in automotive parts and service centres that claim to increase gas mileage, for example, oil and gasoline additives. Your company has been chosen to test one of these claims. Design an investigation to determine the validity of the claim.

14. Research an R-2000 (energy-efficient) home on the Internet, and write a brief description.

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INVESTIGATION 11.1

Designing and Evaluating a Calorimeter

A calorimeter can be as simple as an uninsulated beaker or can. Literature research indicates that very sophisticated calorimeters, such as bomb and flow calorimeters, are also available. There are many laboratory calorimeters between these extremes. In this investigation you will create your own calorimeter. Use modern materials and/or consider what Aboriginal peoples might have used for insulating a calorimeter-like vessel.

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
● Hypothesis	● Procedure	
○ Prediction	● Evidence	

Purpose

The purpose of this investigation is to test a hypothetical design of a calorimeter that is improved (based upon your own scientific and technological criteria) over the simplest calorimeter.

Problem

What is the best design for a simple calorimeter?

Design

Each research group in class uses a different calorimeter design of their choosing. One group uses an uninsulated calorimeter as a control. An individual group and a class Analysis are employed.

Materials

lab apron
eye protection
small soup can
insulation
aluminium foil
duct tape
thermometer or temperature probe
centigram balance
stirring rod
ring and stand
paraffin candle
matches
watch glass
100 mL graduated cylinder
timer

Take care around lit matches and candles. Tie back loose hair and clothing. Never leave a flame unattended.

INVESTIGATION 11.2

Molar Enthalpy of Reaction

Evaluating evidence (dependent upon an experimental design, materials, procedure, and skill) and estimating the certainty of empirically determined values are important skills in interpreting scientific statements. The accuracy (percent difference) of the value for molar enthalpy obtained in this investigation is used to evaluate the calorimeter and the assumptions made in the analysis, rather than to evaluate the prediction and its authority. The ultimate authority in this experiment is considered to be the reference value used in the prediction.

Report Checklist

○ Purpose	● Design	● Analysis
○ Problem	● Materials	● Evaluation (2, 1, 3)
○ Hypothesis	● Procedure	
○ Prediction	● Evidence	

Purpose

The purpose of this investigation is to test the calorimeter design and calorimetry procedure by comparing experimental evidence with a widely accepted value for the molar enthalpy of a neutralization reaction.

Problem

What is the standard molar enthalpy of neutralization for sodium hydroxide when 50 mL of aqueous 1.0 mol/L sodium hydro-xide reacts with an excess quantity of 1.0 mol/L sulfuric acid?

Prediction

The standard molar enthalpy of neutralization for sodium hydroxide is −57 kJ/mol, as per *The CRC Handbook of Chemistry and Physics.*

Wear eye protection. Both sodium hydroxide and sulfuric acid are corrosive chemicals. Rinse with lots of cold water if these chemicals contact your skin.

INVESTIGATION 11.3

Applying Hess' Law

Magnesium burns rapidly, releasing heat and light.

$$Mg(s) + \tfrac{1}{2}O_2(g) \rightarrow MgO(s) \qquad \Delta_c H = ?$$

The enthalpy change of this reaction can be calculated from measurements involving a highly sophisticated calorimeter, but not a polystyrene cup calorimeter. The enthalpy change for the combustion of magnesium can, however, be determined by applying Hess' law to the following three chemical equations.

$$MgO(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2O(l) \qquad \Delta_r H = ?$$

$$Mg(s) + 2\,HCl(aq) \rightarrow MgCl_2(aq) + H_2(g) \qquad \Delta_r H = ?$$

$$H_2(g) + \tfrac{1}{2}O_2(g) \rightarrow H_2O(l) \qquad \Delta_f H° = -285.8\ kJ$$

Purpose

The purpose of this investigation is to use Hess' law to determine a standard molar enthalpy of combustion.

Problem

What is the standard molar enthalpy of combustion for magnesium?

Prediction

The standard molar enthalpy of combustion for magnesium is −601.6 kJ/mol, as per the table of standard molar enthalpies of formation (Appendix I). (The standard molar enthalpy of combustion of magnesium is the same as the standard molar enthalpy of formation of magnesium oxide because both processes have the same chemical reaction equation.)

$$\underset{Mg}{\Delta_c H_m°} = \underset{MgO}{\Delta_f H_m°} = -601.6\ kJ/mol$$

$$Mg(s) + \tfrac{1}{2}O_2(g) \rightarrow MgO(s) \qquad \Delta_c H° = \Delta_f H° = -601.6\ kJ$$

Design

The enthalpy changes for the first two reactions in the series are determined empirically using a polystyrene calorimeter (**Figure 1**), and the third enthalpy change is obtained from Appendix I. The three $\Delta_r H$ values are used, along with Hess' law, to obtain the molar enthalpy of combustion for magnesium.

Report Checklist

○ Purpose	○ Design	● Analysis
○ Problem	○ Materials	● Evaluation (1, 2, 3)
○ Hypothesis	● Procedure	
○ Prediction	● Evidence	

Materials

lab apron
eye protection
magnesium ribbon (maximum 15 cm strip)
magnesium oxide powder (maximum 1.00 g sample)
1.00 mol/L hydrochloric acid (use 50 mL each time)
polystyrene calorimeter with lid and thermometer or temperature probe
50 mL or 100 mL graduated cylinder
laboratory scoop or plastic spoon
steel wool
weighing boat or paper
centigram balance
ruler

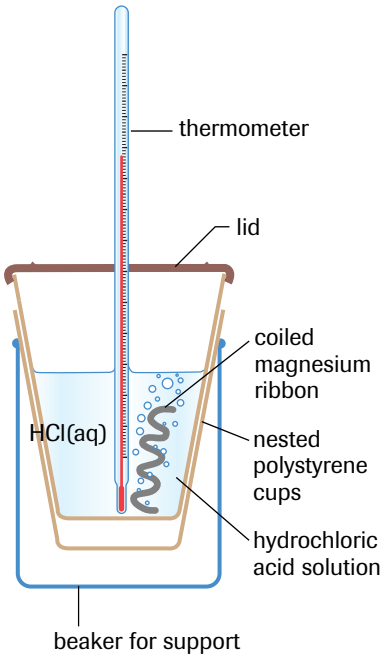

Figure 1
Magnesium ribbon reacts rapidly in dilute hydrochloric acid. With nested polystyrene cups, the enthalpy change can be determined by measuring the temperature change of the hydrochloric acid solution.

Hydrochloric acid is corrosive. Avoid contact with skin, eyes, and clothing. If you spill this acid on your skin, wash immediately with lots of cool water.

Hydrogen gas, produced in the reaction of hydrochloric acid and magnesium, is flammable. Ensure that there is adequate ventilation and that there are no open flames in the classroom.

Magnesium oxide is an extremely fine dust. Do not inhale magnesium oxide powder because it is an irritant.

Chapter 11 SUMMARY

Outcomes

Knowledge

- explain how the sun is a major source of stored chemical energy on Earth (11.1)
- apply the equation $Q = mc\Delta t$ to the analysis of energy transfer (11.2)
- define enthalpy and molar enthalpy for chemical reactions (11.2, 11.3)
- use calorimetry evidence to determine enthalpy changes in chemical reactions (11.2, 11.3, 11.4, 11.5)
- describe photosynthesis, cellular respiration, and hydrocarbon combustion reactions, and understand that combustion and cellular respiration are similar and the reverse of photosynthesis (11.3)
- write and interpret balanced chemical equations incorporating $\Delta_r H$ notation (11.3)
- classify chemical reactions as endothermic or exothermic (11.3, 11.4, 11.5)
- explain and use Hess' law to calculate energy changes for a net reaction from a series of reactions (11.4)
- predict the standard enthalpy change for chemical equations using standard molar enthalpies of formation (11.5)

STS

- state that a goal of technology is to solve practical problems (11.1, 11.2)
- recognize that solving technological problems may require various solutions and have both intended and unintended consequences (11.1, 11.3, 11.4)

Skills

- initiating and planning: design a method to compare molar enthalpy changes when burning fuels (11.2); describe procedures for safe handling, storage, and disposal of materials used in the laboratory, with reference to WHMIS and consumer product labelling information (11.2, 11.4)
- performing and recording: perform calorimetry experiments to determine molar enthalpy changes and use thermometers appropriately (11.2, 11.4)
- analyzing and interpreting: compare energy changes by analyzing data and energy diagrams (11.3, 11.4, 11.5)
- communication and teamwork: work collaboratively using appropriate notation and units for enthalpy changes and molar enthalpies (11.2, 11.3, 11.4, 11.5)

Key Terms

11.2

thermochemistry
isolated system
calorimetry
calorimeter
heat
thermal energy, Q
specific heat capacity, c
joule (J)
enthalpy change, ΔH or $\Delta_r H$
exothermic reaction
endothermic reaction
enthalpy of reaction
molar enthalpy of reaction, $\Delta_r H_m$

11.3

standard molar enthalpy of reaction, $\Delta_r H_m°$
chemical potential energy diagram

11.4

Hess' law

11.5

reference energy state
standard enthalpy of formation $\Delta_f H°$
thermal stability

Key Equations

$Q = mc\Delta t$

$\Delta_r H = n\Delta_r H_m$

$n\Delta H = mc\Delta t$ (calorimetry)

$\Delta_r H = \Sigma\Delta_r H$ (Hess' law)

$\Delta_r H° = \Sigma n\Delta_{fp} H_m° - \Sigma n\Delta_{f\,R} H_m°$ (enthalpies of formation)

MAKE a summary

1. Prepare a table to summarize the four methods for communicating enthalpies of reaction, the quantities and symbols used, the units used. Prepare a second table showing the three methods for determining enthalpy changes. Use the adjectives "molar" and "standard" (and symbols) where appropriate.
2. Refer back to your answers to the Starting Points questions at the beginning of this chapter. How has your thinking changed?

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The following components are available on the Nelson Web site. Follow the links for *Nelson Chemistry Alberta 20–30*.

- an interactive Self Quiz for Chapter 11
- additional Diploma Exam-style Review questions
- Illustrated Glossary
- additional IB-related material

There is more information on the Web site wherever you see the Go icon in this chapter.

EXTENSION

CBC radioONE QUIRKS & QUARKS

A New Kind of Rocket Science

Scientists are looking beyond chemical reactions for rocket propulsion systems. They are investigating nuclear thermal rockets, nuclear electricity, and even solar sailing.

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Many of these questions are in the style of the Diploma Exam. You will find guidance for writing Diploma Exams in Appendix H. Exam study tips and test-taking suggestions are on the Nelson Web site. Science Directing Words used in Diploma Exams are in bold type.

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DO NOT WRITE IN THIS TEXTBOOK.

Part 1

1. A source of energy that is an alternative to fossil fuels is
A. coal
B. solar
C. kerosene
D. natural gas

2. NR Over 95% of the electricity in Alberta is generated by burning fossil fuels. In thermal electric generating stations, energy moves through several forms. List these forms in order.
1 chemical
2 electrical
3 mechanical
4 thermal
___ ___ ___ ___

3. The quantity of thermal energy required to raise the temperature of a unit mass of a substance by one degree Celsius is referred to as
A. specific heat
B. heat capacity
C. specific heat capacity
D. volumetric heat capacity

4. NR A propane camp stove is used to heat 1.50 L of water for making tea. If the water is heated from 5.5 °C to 98.6 °C, the energy absorbed by the water is _____ kJ.

5. Enthalpy change (ΔH) is defined as the change in the
A. thermal energy in a system
B. property of a substance or system that relates to its ability to do work
C. state of a closed system in which all measurable properties are constant
D. chemical energy of a system when the pressure of the system is held constant

6. NR Some coal deposits contain significant quantities of sulfur, which forms sulfur dioxide when the coal is burned. Sulfur dioxide reacts with oxygen in the atmosphere to produce sulfur trioxide, which in turn reacts with rainwater to produce sulfuric acid. Use the following equation to determine the molar enthalpy of reaction for sulfur dioxide.
$2\,SO_2(g) + O_2(g) \rightarrow 2\,SO_3(g) + 197.8\ kJ$
$-$_____ kJ/mol

7. NR The correct reaction equation for the formation of water vapour is
A. $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g) + 120.9\ kJ$
B. $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g) + 241.8\ kJ$
C. $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g) + 483.6\ kJ$
D. $2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(g) + 571.6\ kJ$

8. The equation that correctly represents the enthalpy change involved in the simple decomposition of dinitrogen tetroxide, $N_2O_4(g)$, is
A. $N_2(g) + 2\,O_2(g) \rightarrow N_2O_4(g)$ $\Delta_r H = +11.1$ kJ
B. $N_2(g) + 2\,O_2(g) \rightarrow N_2O_4(g)$ $\Delta_r H = -11.1$ kJ
C. $N_2O_4(g) \rightarrow N_2(g) + 2\,O_2(g)$ $\Delta_r H = +11.1$ kJ
D. $N_2O_4(g) \rightarrow N_2(g) + 2\,O_2(g)$ $\Delta_r H = -11.1$ kJ

Use the following statements to answer questions 9 and 10.
1. Carbon dioxide is a reactant.
2. Carbon dioxide is a product.
3. Water is a reactant.
4. Water is a product.
5. The reaction is exothermic.
6. The reaction is endothermic.

9. NR Which of the above statements are true for photosynthesis?
___ ___ ___ ___

10. NR Which of the above statements are true for cellular respiration?
___ ___ ___ ___

Consider these fuels when answering questions 11 to 13.
1. methane
2. propane
3. octane
4. methanol

11. NR List the fuels in order of increasing standard molar enthalpy of formation.
___ ___ ___ ___

12. NR List the fuels in order of increasing standard molar enthalpy of combustion.
___ ___ ___ ___

13. NR List the fuels in order of increasing thermal stability.
___ ___ ___ ___

14. NR The standard enthalpy change for the world-scale catalytic cracking of ethane into ethene, for example, near Red Deer and near Fort Saskatchewan, Alberta, is
+_____ kJ

Part 2

15. **Explain**, in general terms, how stored energy in the bonds of hydrocarbons originated from the sun.

16. **Distinguish** between an enthalpy change and a molar enthalpy including general symbols and word descriptions.

17. DE The energy content of foods can be determined by combustion analysis using a calorimeter. The combustion of 1.25 g of peanut oil caused the temperature of 2.0 kg of water to increase by 5.3 °C.

(a) **Determine** the increase in thermal energy of the water.

(b) **Determine** the energy content of peanut oil in kilojoules per gram.

18. DE Methylpropane, $C_4H_{10}(g)$, is used as a lighter fluid. When 1.52 g of methylpropane is burned in a calorimeter, the temperature of 2.23 kg of water changed by 8.04 °C.

(a) Calculate the change in thermal energy of the water.

(b) **Determine** the enthalpy change for the combustion of 1.52 g of methylpropane under these conditions.

(c) If you were using this evidence to report the molar enthalpy of combustion for methylpropane, what symbol, value, and sign would you use? **Justify** your answer.

19. The enthalpy of neutralization of an acid–base reaction can be determined using a graduated cylinder, a thermometer, and a simple polystyrene calorimeter. Complete the Analysis of the investigation report.

Problem

What is the molar enthalpy of neutralization for ethanoic acid?

Design

Measured volumes of ethanoic acid and sodium hydroxide are mixed in a polystyrene calorimeter. The temperatures of the solutions are measured before and after mixing.

Evidence

volume of 1.00 mol/L ethanoic acid	= 50.0 mL
volume of 1.00 mol/L sodium hydroxide	= 50.0 mL
initial temperature of ethanoic acid	= 22.5 °C
initial temperature of sodium hydroxide	= 22.5 °C
final temperature of mixture	= 27.8 °C

20. DE As an alternative to chemical energy, heating water is one of the most cost-effective uses of solar energy, providing hot water for showers, dishwashers, and washing machines.

(a) **Determine** the thermal energy acquired (and chemical energy saved) when the temperature of 100 L of water is raised from 6.5 °C to 58.5 °C.

(b) What mass of natural gas (methane) would have to be burned to heat the same volume of water by the same temperature change?

(c) A domestic solar water heater saves 32.5 MJ of chemical (combustion) energy while increasing the thermal energy of some water from 12.3 °C to 45.5 °C. What mass and volume of water did the water heater contain?

21. DE Prepare a general comparison of photosynthesis, cellular respiration, and hydrocarbon combustion. Your response, which could be a paper or electronic graphic or table, should include

- reactants
- products
- endothermic/exothermic description
- sign of the enthalpy change

22. The flame of an oxyacetylene torch is hot enough to melt most metals. Calculate and then communicate the energy changes for the combustion of acetylene (ethyne) in four different ways.

23. Although calorimetry is the basic method of obtaining enthalpy changes, this method cannot always be directly used for many reactions. **Describe** two other methods commonly used to obtain an enthalpy change for a reaction. Include what information is required, where this information is obtained, and any mathematical equations.

24. Vinyl chloride monomer is produced at Fort Saskatchewan, Alberta, to produce, in turn, polyvinyl chloride (PVC). Use Hess' law to **determine** the enthalpy of reaction for the conversion of ethene into chloroethene (vinyl chloride).

$$C_2H_4(g) + Cl_2(g) \rightarrow C_2H_3Cl(g) + HCl(g) \qquad \Delta_r H = ?$$

$$C_2H_4(g) + Cl_2(g) \rightarrow CH_2ClCH_2Cl(l) \qquad \Delta_r H = -179.3 \text{ kJ}$$

$$C_2H_3Cl(g) + HCl(g) \rightarrow CH_2ClCH_2Cl(l) \qquad \Delta_r H = -71.9 \text{ kJ}$$

25. To meet the demand for gasoline, hydrocarbons of large molar mass are broken into smaller fragments in a process called *cracking*. An overall reaction equation for producing octane from hexadecane is

$$C_{16}H_{34}(l) + H_2(g) \rightarrow 2\,C_8H_{18}(l) \qquad \Delta_r H^\circ = ?$$

Use the following reaction equations to **determine** the standard enthalpy change for the above reaction.

$$2\,C_{16}H_{34}(l) + 49\,O_2(g) \rightarrow 32\,CO_2(g) + 34\,H_2O(l) \qquad \Delta_r H^\circ = -21\,446 \text{ kJ}$$

$$2\,C_8H_{18}(l) + 25\,O_2(g) \rightarrow 16\,CO_2(g) + 18\,H_2O(l) \qquad \Delta_r H^\circ = -10\,940 \text{ kJ}$$

$$2\,H_2(g) + O_2(g) \rightarrow 2\,H_2O(l) \qquad \Delta_r H^\circ = -572 \text{ kJ}$$

26. DE In the metal refining industry, sulfide ores are usually roasted in air to produce metal oxides and sulfur dioxide. For example, copper(II) sulfide is converted to copper(II) oxide as part of the refining of copper.

(a) Use referenced data to **determine** the standard molar enthalpy of roasting for copper(II) sulfide.

(b) Draw a chemical potential energy diagram for this roasting reaction.

27. Calculate the standard enthalpy of reaction for the following organic reactions in the Alberta chemical industry studied in Unit 5.

(a) octane reacts with hydrogen (i.e., undergoes hydrocracking) to produce pentane and propane

(b) ethene reacts with benzene to produce phenylethene (polystyrene monomer) and hydrogen

28. The following investigation was carried out to find the enthalpy of formation, $\Delta_f H$, for calcium oxide. Complete the Prediction and the Analysis of the investigation report.

Problem

What is the enthalpy change for the reaction?

$$Ca(s) + \frac{1}{2} O_2(g) \rightarrow CaO(s)$$

Design

Calcium metal reacts with hydrochloric acid in a calorimeter and the enthalpy change is determined. Similarly, the enthalpy change for the reaction of calcium oxide with hydrochloric acid is determined. These two chemical equations are combined with the formation equation for water to determine the required enthalpy change.

Evidence

concentration of HCl(aq) = 1.0 mol/L
volume of HCl(aq) = 100 mL

Table 1

Reactant	Mass (g)	Initial temperature (°C)	Final temperature (°C)
Ca(s)	0.52	21.3	34.5
CaO(s)	1.47	21.1	28.0

29. Cars and furnaces often do not run at their highest energy efficiency.

(a) Calculate the standard molar enthalpy of complete combustion of propane (to produce liquid water).

(b) Calculate the standard molar enthalpy of complete combustion of propane (to produce water vapour).

(c) Calculate the standard molar enthalpy of incomplete combustion for propane. Use the following chemical equation for the incomplete combustion of propane.

$$2\,C_3H_8(g) + 8\,O_2(g) \rightarrow C(s) + 2\,CO(g) + 3\,CO_2(g) + 8\,H_2O(g)$$

(d) List the above standard molar enthalpies of combustion of propane from lowest to highest.

(e) Research the features of a high-efficiency (condensing) furnace, and relate these features to your calculations above.

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30. According to an Inuit Elder, the refrigerator has damaged the Inuit community and lifestyle because families do not have to share in the same way that they used to. Suggest possible community interactions before and after the arrival of refrigerators. What does this imply about our notion of technological progress?

Extension

31. Choose an "alternative" fuel to research. Select and integrate information (including graphs, tables, and graphics) from several sources. Your information should illustrate at least three perspectives on the use of this fuel. Decide on the target audience, and present your findings attractively and persuasively.

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32. Thermal-electric power stations burn coal (various types), fuel oil (heavy and light), natural gas, or wood to produce steam to drive a turbine and generator. The efficiency of converting the available chemical energy into electrical energy can be expressed as a percentage. The Statistics Canada publication *Human Activity and the Environment* provides this kind of data shown in **Table 2**.

(a) **Analyze** the data in Table 2 (or obtain more recent data) to find any trends or other insights.

(b) **Identify** other non-combustible energy sources that can be used to produce steam in a power station.

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Table 2 Percent Efficiency of Thermal-Electric Power Stations by Fuel Type (%)

	Coal					Fuel oil		Gas	Wood
Year	CB	IB	CSB	ISB	Lig	HFO	LFO	NG	W
1998	34.01	37.21	32.51	34.91	29.71	33.01	32.21	34.71	40.01
1999	33.91	35.61	32.71	35.11	29.11	33.21	30.41	35.41	35.41
2000	31.31	34.81	32.81	33.21	30.11	33.01	28.31	35.11	31.31
2001	32.41	32.81	32.41	28.51	30.11	33.01	29.31	34.41	26.81
2002	30.21	33.91	32.81	33.01	30.21	33.51	29.41	35.91	27.11

CB: Canadian bituminous; IB: imported bituminous; CSB: Canadian sub-bituminous; ISB: imported sub-bituminous; Lig: lignite; HFO: heavy fuel oil; LFO: light fuel oil and diesel; NG: natural gas; W: wood